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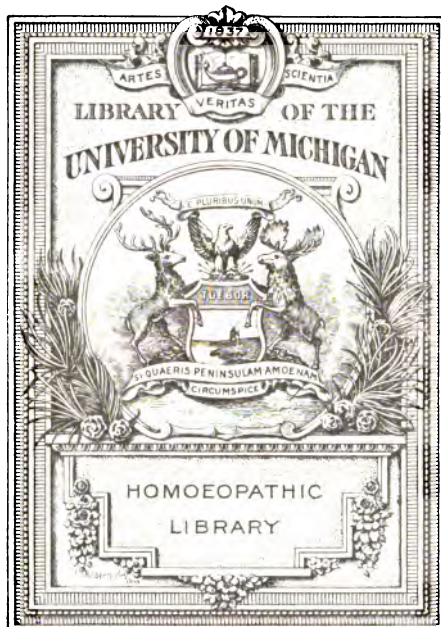
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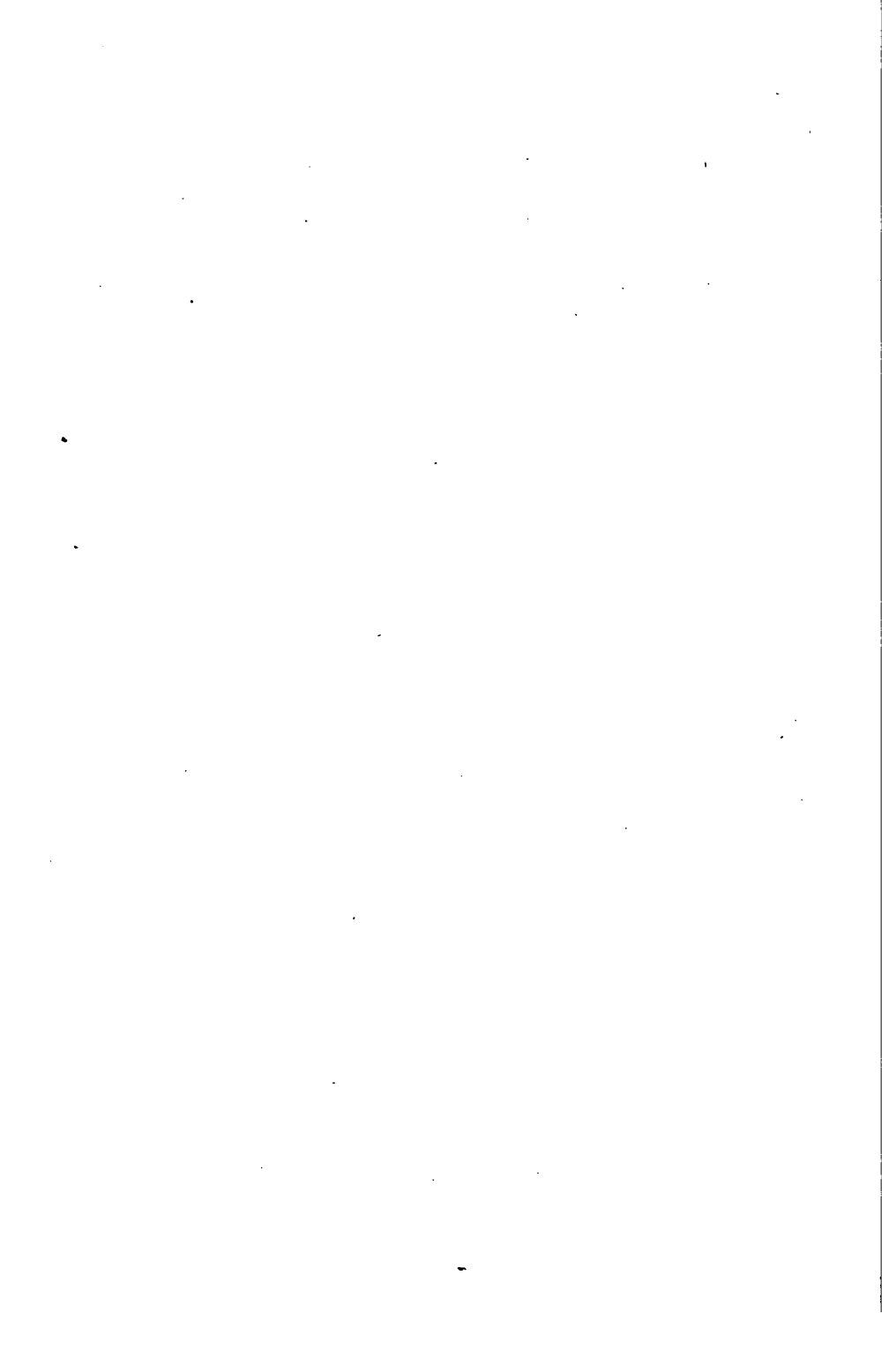
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A TEXT-BOOK  
OF  
CHEMISTRY AND CHEMICAL  
URANALYSIS FOR NURSES

BY

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## PREFACE.

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THE demand for a simple book on chemistry, written especially for, and adapted to, the needs of the nurse, has become more and more urgent with the institution of regular courses on chemistry in the best training schools. In the preparation of this work the author has endeavored to cover the subject briefly and clearly, and in an interesting style, so that the reader may easily absorb and assimilate the material presented.

It is not to be questioned that the more chemistry a nurse knows in usable form the greater her value to the patient and to the physician. In many cases the important measure of feeding the sick is left to the nurse. In most instances her empirical knowledge of dietetics may guide her aright. Yet it is obvious that some knowledge of the chemical composition of food-stuffs and of the chemical processes of digestion and assimilation will better serve her in cases of unusual type and those presenting metabolic and digestive disturbance. If the nurse knows no chemistry, how can she be expected always to remember that starches yield sugars and are to be given to the diabetic with extreme caution? What can the term Calorie mean to her? There are other questions, too, of drug admin-

istration and application, of the collection and preservation of specimens, and of discerning observation.

It is the author's impression that chemistry has not yet received the attention in undergraduate and graduate instruction of nurses that its importance merits. He believes that one of the chief reasons is that few teachers have the time to collect and put into simple form the material necessary for the nurse's foundation and understanding in chemistry.

In this volume the author has striven for simplicity and a gradual logical development of the subject; and he hopes that the book will adequately fill the place in nursing education for which it was designed.

H. L. A.

NEW YORK, 1915.

# CONTENTS.

---

## CHAPTER I.

THE SLAKING OF LIME . . . . .	19
-------------------------------	----

## CHAPTER II.

WEIGHTS AND MEASURES . . . . .	24
--------------------------------	----

## CHAPTER III.

THE METALS . . . . .	25
----------------------	----

## CHAPTER IV.

MOLECULES AND ATOMS . . . . .	30
-------------------------------	----

## CHAPTER V.

CHEMICAL PROCESSES . . . . .	34
------------------------------	----

## CHAPTER VI.

ATOMIC WEIGHTS . . . . .	39
--------------------------	----

## CHAPTER VII.

OXYGEN . . . . .	43
------------------	----

## CHAPTER VIII.

HYDROGEN . . . . .	48
--------------------	----

## CHAPTER IX.

WATER . . . . .	5
-----------------	---

## CHAPTER X.

HEAT . . . . .	60
----------------	----

## CHAPTER XI.

SOLUTIONS AND PURIFICATION OF SUBSTANCES . . .	64
--	----

## CHAPTER XII.

NATURAL WATERS—CHEMICAL ACTION OF WATER . .	68
---	----

## CHAPTER XIII.

COMPOSITION OF WATER . . . . .	72
--------------------------------	----

## CHAPTER XIV.

HYDROGEN PEROXIDE . . . . .	77
-----------------------------	----

## CHAPTER XV.

CHLORINE . . . . .	81
--------------------	----

## CHAPTER XVI.

BROMINE—IODINE—FLUORINE . . . . .	89
-----------------------------------	----

## CHAPTER XVII.

SULPHUR . . . . .	95
-------------------	----

## CHAPTER XVIII.

SODIUM . . . . .	100
------------------	-----

## CHAPTER XIX.

ACIDS AND BASES—POTASSIUM . . . . .	106
-------------------------------------	-----

## CHAPTER XX.

PHOSPHORUS—ARSENIC—ANTIMONY—BISMUTH . . .	112
---	-----

## CONTENTS

vii

### CHAPTER XXI.

CALCIUM . . . . .	119
-------------------	-----

### CHAPTER XXII.

MAGNESIUM GROUP . . . . .	123
---------------------------	-----

### CHAPTER XXIII.

ALUMINUM—IRON—MANGANESE . . . . .	129
-----------------------------------	-----

### CHAPTER XXIV.

LEAD—SILVER—PLATINUM . . . . .	133
--------------------------------	-----

### CHAPTER XXV.

CARBON . . . . .	138
------------------	-----

### CHAPTER XXVI.

COMPOUNDS OF CARBON WITH HYDROGEN . . . . .	141
---	-----

### CHAPTER XXVII.

ETHERS . . . . .	150
------------------	-----

### CHAPTER XXVIII.

THE MARSH GAS SERIES . . . . .	153
--------------------------------	-----

### CHAPTER XXIX.

THE PARAFFINS . . . . .	160
-------------------------	-----

### CHAPTER XXX.

SUGARS . . . . .	165
------------------	-----

### CHAPTER XXXI.

POLYSACCHARIDS . . . . .	173
--------------------------	-----

## CHAPTER XXXII.

THE DIGESTION OF CARBOHYDRATES . . . . .	180
--	-----

## CHAPTER XXXIII.

FATS . . . . .	183
----------------	-----

## CHAPTER XXXIV.

BENZENE SERIES . . . . .	191
--------------------------	-----

## CHAPTER XXXV.

NITROGEN . . . . .	202
--------------------	-----

## CHAPTER XXXVI.

OTHER NITROGEN COMPOUNDS . . . . .	208
------------------------------------	-----

## CHAPTER XXXVII.

PROTEINS . . . . .	216
--------------------	-----

## CHAPTER XXXVIII.

THE BLOOD . . . . .	226
---------------------	-----

## CHAPTER XXXIX.

MILK . . . . .	235
----------------	-----

## CHAPTER XL.

THE URINE . . . . .	240
---------------------	-----

## CHAPTER XLI.

URANALYSIS . . . . .	249
----------------------	-----

## TO THE INSTRUCTOR.

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MANY students of medicine and nurses find the particular chemistry which they need very difficult to acquire. Obviously, the manner of presentation of the subject will affect largely the results, especially in the training school where the all-important demonstration facilities are usually lacking and where there is relatively little time devoted to the teaching of chemistry. It is exceedingly difficult to really grasp this subject by reading or listening to lectures unaccompanied by experimental work, and the student is soon found simply attempting to memorize a large number of isolated facts. Especially is this true in the beginning, and chemistry soon becomes a hopeless muddle. Demonstrations, therefore, are urgently recommended, and whenever possible even a small number of laboratory exercises should be arranged so that the student can handle the test-tube and chemicals. In this manner it is often possible to make the subject fascinating and interesting instead of drudgery. This compilation is necessarily limited to those subjects which the efficient nurse must know: many processes of which all educated persons should possess at least a little knowledge have been omitted. If the student acquires



fundamental ideas of chemistry these things will come easily later in their general reading.

However, it is not to be expected that the average student will digest all the matter contained herein. Enough has been left in the reduction of the original manuscript to satisfy the interested student and not too much to frighten the uninterested.

An attempt is made to develop the subject from the simplest and more familiar phenomena. Since the interest of the author was first aroused in this subject by the slaking of lime this is made the foundation for the introduction of the nurse to this important science.

# CHEMISTRY FOR NURSES.

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## CHAPTER I.

### THE SLAKING OF LIME.

WHEN water is poured over quicklime, heat is generated and the water soon begins to boil. If water is poured on the same amount of limestone from which the lime is made we obtain no change in temperature and no heat is generated; hence, there must be some change other than the physical contact of the lime and water in order to produce this large amount of heat. If there is such a change, what does it amount to and how is it produced? Quicklime is made by heating very intensely and over a rather long period of time the stones we know as limestone rock. If we heat absolutely dry limestone in the kiln for some hours we find by reweighing it from time to time that it gradually loses weight. Continuing the heating processes we find that we arrive at a stage where it no longer loses weight, that is, we have heated it to *constant weight*, and no amount of heating can reduce the weight further. Curiously enough, the loss is always 44 pounds in a hundred. What does it mean then if we take several samples of pure dry

limestone and heat it in the kiln to constant weight and find that exactly 44 out of every 100 pounds is lost? Can we not say, first, that the limestone probably has a definite composition—that is, it is composed of certain definite materials, and second that these materials whatever they are occur in the same proportion each time?

During this process of heating, when the limestone is being transformed into lime, the physical characters of the limestone undergo marked changes, chief among them is the change from a hard, granular substance to a smooth, friable, soft material. But what happens during the heating? It will be shown later, that this 44 pounds out of a hundred which is given off is the same gas that we give off constantly from our lungs, namely, carbon dioxide. And the driving off of this gas by heat is a chemical process and a chemical change is brought about. We could grind the original limestone into a very fine powder—that would be a physical change—yet on pouring water over it we would get no heat. If on the other hand we raise this powder to red heat for several hours, we have driven off a part of it though we can not see it come off and have produced a change in the internal structure of the substance (chemical change). We also witness a chemical change when water is added to the lime. Suppose that we add to our 56 pounds of lime resulting from the heating of 100 pounds of limestone, a hundred pounds of water at room temperature (68° F.) and stir for a few minutes—always the temperature of the water increases the same number

of degrees it makes no difference how many experiments you make. Then again there must be some definite and always constant change taking place when water is poured on lime. Now if we try to recover our hundred pounds of water by distillation (a physical process) we find that at the temperature of boiling water (212° F.), we can recover only 82 pounds of the 100 pounds of water which we added. Therefore 18 pounds of water must have been fixed to the quicklime in the slaking process. There must have been some chemical change also for our lime is no longer a smooth non-crystallizing substance but consists of clear crystals having definite angles. Since water or some of the constituents of water has been added to our quicklime we can call this new crystallized substance *hydrated* lime. This new product weighs 74 pounds and does not give off water when heated to boiling point. If now we heat our dry hydrated lime to a higher degree it begins to lose weight and water is given off until finally the whole of the 18 pounds of water is recovered and we have 56 pounds of quicklime left, having properties identical with those before described. This is another chemical change—the water was driven off by the heat leaving behind a compound with different physical properties.

Suppose we add again to our 56 pounds of quicklime 100 pounds of water. When the solution has cooled and settled clear, remove a small portion in a test-tube and blow through it by means of a glass tube. The clear solution becomes turbid and a white precipitate is formed. Suppose we bubble by means of a blower

the used air which collects in the top of a theatre or the gases given off from a lime kiln, through our mixture of 56 pounds of quicklime and 100 pounds of water. Finally there comes a stage when our breath bubbled through the clear solution no longer causes turbidity. Now should we distill the water and weigh it, provided none has been accidentally lost in the operation we find that it weighs 100 pounds. Obviously something has replaced the 18 pounds of water which in the first instance was held in chemical combination by the quicklime. Furthermore, our white residue now weighs 100 pounds and chemical tests show us that this substance is identical with the limestone which we started with. We know that the carbon dioxide in the expired breath has combined with the hydrated lime to form limestone. Then we can now call limestone carbonated lime. These are examples of chemical changes. We have learned that no chemical change takes place without a corresponding physical change and also that these chemical compounds unite in definite proportions. Having once learned by experiment how much of one substance is held in chemical union with another we are able to predict before we put these substances together just what will happen and to calculate how much substance we shall have at the end of our experiment.

#### SUMMARY OF CHAPTER I.

A *physical* change means a change in form, as the grinding of limestone. The freezing of water, and its conversion into steam are examples of physical changes.

In a physical change the form or state of a substance is changed while the intrinsic chemical properties are not altered.

A *chemical* change means a change in composition. It was shown that the heating of limestone causes a chemical change: a gas (carbon dioxide) is given off and the substance left is lime which has physical and chemical properties totally different from limestone. When water is added to lime a chemical change takes place; a new compound possessing new properties is formed. Heat is given off. In all chemical changes heat is either absorbed or given off.

In any chemical change there is always a physical change; the reverse is never true.

Pure chemical substances contain definite elements in definite proportions.

## CHAPTER II.

### WEIGHTS AND MEASURES.

IN our experiments with quicklime and water we used such quantities stated in such terms as we can readily appreciate. As a matter of fact in order to obtain absolutely accurate results we must use very small quantities. We used 100 pounds so that we should not be confused by decimals. In chemical work the metric system is used in all countries. One *kilogram* (abbreviation *kilo* or *k*) is equivalent to 2.2 pounds. One thousandth of a *kilogram*, a *gram*, is used as the basis of the metric system; about 454 *grams* equal 1 *pound*. An ounce avoirdupois equals about 28.5 *grams*. Fifteen to twenty drops of pure water weigh 1 *gram*, and the volume occupied by this amount of water is equal to 1 *cubic centimeter* (1 c.c.). *Decigram*, one tenth of a gram, is rarely used, but milligram 0.001 gm. is convenient for stating the doses of the more active medicinal agents. Sixty-five milligrams (0.065 gm.) are equivalent to 1 grain.

## CHAPTER III.

### THE METALS.

A CLASS of substances represented by gold, silver, copper, tin, lead, iron, nickel, quicksilver, etc., possessing the power of conducting heat and electricity, capable of being fused, moulded and of being drawn out into various shapes, and having a peculiar luster, are known to us as metals. Some of them (gold, silver, lead, copper, quicksilver, etc.), occur in comparatively large quantities as metals in nature and they also occur in combinations with other substances as ores. With one exception (quicksilver) at ordinary temperatures they are solid.

**Rusting of the Metals.**—Our most common example of a metal is iron. This metal is found free in nature in very small quantities. Pure iron has a grayish-white color which on exposure to damp air soon changes to black and then to red. In the process of color change, physical properties of the metal change: if this entire piece is rusted throughout it loses its strength and crumbles, it can no longer conduct electricity or be hammered into various shapes. Its weight increases. The rusting process can be hastened by heating in the air.

**The Air and the Rusting Process.**—A piece of iron placed in a glass tube, the air exhausted and the tube



sealed, will retain forever its luster. Even heat will not bring about the tarnishing or rusting process unless air is admitted. Therefore air or some constituent of air is necessary for rusting.

Silver will rust quickly, turning black, when heated in air. Let us, therefore, place some silver filings in a hard glass tube and heat it to redness and allow air to pass through slowly and then to a similar tube containing pieces of untarnished iron. So long as the silver continues to rust the iron will remain untarnished. Conduct the air, which has passed over the heated silver and iron to a bell jar and when it is collected in sufficient quantities place in it a mouse. The mouse dies very quickly. Now strike a match and thrust it into the bell jar—the flame goes out. What have we learned? First, that the substance in the air which causes iron to rust is the same as that which brings about the tarnishing (rusting) of silver; second, that this same substance is also necessary for life; third, that it is also necessary for *burning*. Therefore, we reason that the rusting of metals, the maintenance of life by respiration and the burning of wood are closely related if not the same processes.

**Oxygen.**—To continue our experiment, let us heat some of the rusted silver and collect the gas given off in the process by conducting it into a bell jar. Now repeat the mouse and the match experiments with this new gas and we get very different results. The mouse is enlivened and the match burns more brightly than in ordinary air. These results lead us to believe that we have gotten that substance in the

air essential for life in a pure state. Other experiments show us that this is true. Priestly, an English chemist, who afterward came to America and settled at Northumberland, Pa., discovered these facts by working with rusted quicksilver and called the gas thus obtained *dephlogisticated air*. Lavoisier later called it *oxygen*, which means *acid producing*, thinking (erroneously) that this substance always produces acid.

Now we term the rusting of the metals, the burning of substances and the various vital processes in the animal body where oxygen is combined with other substances, *oxidations*. The new chemical compounds thus produced are known as the *oxides*. Iron rust is chemically known as *iron oxide*, etc.

**Elements.**—From the foregoing we have learned that there are at least two gases in the air which may be easily taken out, namely, oxygen and carbon dioxide. We know that these substances exist there free or in other words air is a *mixture of gases*. As has been stated, the oxides and other chemical compounds of metals occur in the earth as ores. Here we have examples of both compounds and mixtures; the oxide of iron is a compound consisting of iron and oxygen in chemical union, whereas it may be *mixed* with clay or other oxides that may be removed mechanically. Then a mixture may be analyzed mechanically but compounds can be separated into their constituents by *chemical* processes only.

In the process of chemical analysis we arrive at a stage where we can no longer separate the various

substances into simpler materials. Such a substance, which can no longer be resolved into unlike components, we term an *element*. The metals already mentioned are elements. Brass is an alloy or mixture of two elements, copper and zinc. Oxygen is also an element.

**Symbols.**—For simplicity, convenience and economy in writing chemical formula, the names of the various elements are abbreviated or the first letter alone or combined with some other distinctive letter is used; for example, O means oxygen and Os osmium; H means hydrogen and He helium; S sulphur and Si silicon. We hardly see how Fe stands for iron or Cu for copper until we remember the Latin names *ferrum* and *cuprum*. Then if an oxide of iron is composed of 1 atom of iron and 1 atom of oxygen we would write it thus FeO; if in another oxide of iron there are two atoms of iron and three atoms of oxygen we write it Fe<sub>2</sub>O<sub>3</sub>.

### SUMMARY OF CHAPTER III.

All substances can be analyzed into their simple components which are known as *elements*. Metals (iron, gold, silver, copper and mercury) are elements. They cannot be further separated into different substances. Oxygen is also an element. It is a gas occurring in the air and is necessary to life. The rusting of metals is a chemical union of the element oxygen with the metals (also elements). The burning of wood is also an oxidation. Therefore oxidation may take place quickly resulting in the production of heat and

a flame or it may proceed quietly and slowly and in the presence of water. In either case oxidation is the same and the total amount of heat given off is exactly the same.

In writing chemical reactions symbols are made use of to indicate the various elements. Usually the first letter, capitalized, is used as the symbol for the element, O=oxygen, H=hydrogen. Sometimes the first two letters, as Si=silicon (S=sulphur). Cu=copper (Latin, cuprum). Ag=silver (Latin, argentum).

## CHAPTER IV.

### MOLECULES AND ATOMS.

THE smallest particle of a compound which can exist is called a *molecule*. The size of molecules vary but they are inconceivably small. Since compounds are made up of two or more elements, then molecules may be still further divided into *atoms*, which according to Dalton's theory are the indivisible particles of which all substances are composed. If we take FeO as an example, the smallest particle of FeO which can exist is a *molecule*, and this *molecule* consists of an Fe atom and an O atom, which are indivisible. Generally *atoms* do not exist alone but only in combination, that is to say an element for example like oxygen consists of *molecules* of oxygen each of which is composed of two *atoms* of oxygen. We do not have then simply O in the air but O<sub>2</sub>, for when an atom of O has no Fe or Cu or other element to combine with, it unites with another atom of O making O<sub>2</sub>.

The size of an atom is so small that it cannot be determined or even imagined. We know that certain substances give off odors for long periods of time and never diminish in weight. In an attempt to convey some idea as to the size of these minute, infinitesimal particles, Lord Kelvin says: "Imagine a rain drop or a globe of glass as large as a pea, to be magnified up

to the size of the earth; each constituent being magnified in the same proportion. The magnified structure would be coarser grained than a heap of small shot, but probably less coarse grained than a heap of cricket balls."

**Can Matter be Destroyed?**—In our experiments with limestone, we certainly changed this chemical compound and reduced its weight by heating. It was found that we drove off a gas and left a white friable substance that would unite with water and in the process would still further suffer chemical and physical changes. But we found that after we obtained this new chemical compound, hydrated lime, we could bubble the gases from the kiln through it and obtain not only the same *substance* (limestone) with which we started but in exactly the same amount. In other words with careful handling no matter was lost. Consider another example: iron rusts and increases in weight—and loses its metallic properties. By appropriate chemical processes we are able to drive away the oxygen and recover in exactly the same amount the metallic iron. Well enough, but how about the burning of coal—do we not destroy the coal in this burning process which we have learned to regard as an oxidation? Does not the ash weigh less than the coal did? The ash does weigh less, for this represents only the mineral parts of the coal. If we had taken the trouble to collect every particle of the gases given off in this combustion (burning) we would have found that the weight of the gases plus the weight of the ash exceeds the weight of the

coal and the sum is exactly equal to the weight of the coal plus the weight of the oxygen used in the process. Therefore nothing is destroyed—things may be changed physically and chemically and the products wafted to the four winds of the earth but never destroyed. The ashes may be washed away and the gases dissipated but eventually they are gathered into nature's laboratory and rebuilt into combustible materials. If the elements were destroyed all our methods of determining the exact amounts of each present in a compound would be useless. The fact that in chemical reactions there is no change in the weight (mass) of each element is known as the *law of the conservation of mass*.

**Energy.**—The woodsman's method of lighting a fire is to rub two pieces of wood together until the heat generated is sufficient to start the oxidation of the wood. By means of the friction mechanical energy was transformed into heat and finally when the heat is present in large quantities another form of energy, namely, light, manifests itself. Heat energy may be transformed into mechanical through the agency of the steam engine and then into electrical energy by the turning of a dynamo. Electrical energy may in turn be transformed into light, mechanical energy or heat by means of incandescent filament, motor and coils, respectively.

Just as matter can not be destroyed it is not possible to destroy energy. Energy may be transformed into heat and dissipated (*i. e.*, lost to immediate surroundings), but it always remains energy, and always exists

in the same amount. In the transformation of the energy contained in coal (latent energy) into steam or electricity (potential or kinetic energy) we may not always finish our experiment with the same amount with which we began but this is due to imperfections in our machinery. The energy still exists somewhere either in a *latent* or *kinetic* form. This is known as the *principle of the conservation of energy*.

#### SUMMARY OF CHAPTER IV.

A compound is the result of the union of two elements. The smallest unit of a compound is a molecule. A molecule is composed of atoms. Iron oxide,  $\text{FeO}$ , consists of very small bodies (molecules) of  $\text{FeO}$ , which in turn are made up of Fe (atom) and O (atom). A molecule of an *element* may consist of one, two, three or more atoms, a molecule of a compound consists of at least two atoms.

Matter may suffer chemical or physical change but it cannot be destroyed. In chemical reactions there is no change in the weight (mass) of the substances involved: this is the *law of the conservation of mass*.

Energy undergoes change in form (heat, light, electrical current), but is never destroyed. Heat may be radiated and lost to immediate surroundings but it will always exist as energy in some form.



## CHAPTER V.

### CHEMICAL PROCESSES.

IN raising a weight to a given height a certain amount of energy (kinetic) is required. What becomes of it—the weight isn't heated or lighted or electrically charged in the process. It is simply stored up as latent energy—for when that weight falls and strikes it liberates exactly the same amount of energy that was required to lift it. Now, returning to our limestone experiment, we will remember that in driving off the carbon dioxide it was necessary to use heat; and when we poured water on the quicklime heat was liberated. In the first chemical process heat was absorbed (compare with the raising of the weight), in the second heat was liberated (falling of the stone). In every chemical process heat is either absorbed or liberated. And we have just come to understand why the water boils when it is poured on quicklime. It means that there is a chemical union brought about between the quicklime and water with the liberation of heat. When the lime hydrate is subjected to heat the reverse is true; that is, the water is driven off and heat absorbed in the process. The amount of heat given off or absorbed in any chemical process is always the same for a given mass of substances. Later we shall learn that our bodies are kept warm by the heat which is liberated during chemical processes.

**The Law of Constant Proportions.**—If the amount of heat absorbed or liberated is always the same for a given mass of reacting substances as we have just learned then we presume that these substances react in a constant ratio to one another. As a matter of fact, this observation simply confirms what had already been found out by different methods. In the early development of chemistry various substances were analyzed in a crude way, but the methods were sufficient to detect the constancy with which the elements occurred in a given chemical substance. For instance, in our lime experiment we started with 100 pounds—on heating until constant weight was attained we always had 56 pounds, no more, no less, provided we started with 100 pounds of pure water-free lime. It has been found that whenever quicklime is air slaked, that is, allowed to absorb all the carbon dioxide it will take up, that the combining ratio is always 56 parts of quicklime to 44 parts carbon dioxide. Many other examples can be given to prove that *if two substances react chemically and form a third, they enter into combination in a constant proportion.*

**The Law of Multiple Proportions.**—It has been found that two or three elements may combine to form entirely different substances. Methods for merely detecting the presence of the elements (known in chemistry as Qualitative Analysis) would show no difference between these two compounds, while *quantitative analysis* (*i. e.*, determining the amounts of the elements present) would bring out this difference in composition. Two substances which are used exten-

sively in medicine illustrate this law. Calomel, a white insoluble, non-crystalline, non-poisonous compound, is composed of one atom of mercury (quick-silver) and one atom of chlorine (a greenish, pungent gas and a constituent of common salt). The chemical formula for mercury-chloride is  $\text{HgCl}$ . If instead of *one* atom of chlorine combined with mercury we have *two*, namely  $\text{HgCl}_2$ , an entirely different chemical substance results. The latter substance is *bichloride* of mercury or corrosive sublimate, a soluble, highly poisonous, crystalline compound, though it contains nothing that is not found in calomel. The difference lies in the relative amounts of each. By quantitative analysis we observe that the number of atoms of chlorine in the second compound is exactly twice the number in the first for every atom of mercury. There are no compounds of mercury and chlorine known in which the number of atoms of chlorine is one and a half or three-quarters, etc., times the number found in calomel. There are only two compounds of mercury and chlorine known: if no law governed their combination there would be any number of different substances depending upon the relative amounts of each present at the beginning of the reaction. It has been found that when *two elements combine in more than one proportion, the masses of the one which combine with a given mass of the other bear a simple rational relation to one another*. This ratio in the mercury compounds is 1 to 2. When we come to study the oxides of iron, for example, we find several compounds possible:  $\text{FeO}$ ,  $\text{Fe}_3\text{O}_4$ ,  $\text{Fe}_2\text{O}_3$ . These com-

pounds may at first glance seem at variance with our law and our O seems to be present in fractional parts, but let us elevate them to a common iron content as  $\text{Fe}_2$ , then our compounds become  $\text{Fe}_6\text{O}$ ,  $\text{Fe}_6\text{O}_8$  and  $\text{Fe}_6\text{O}_9$ . The combining ratios of the O is 6, 8 and 9. When we come to study the various combinations of carbon and hydrogen, we find a wider application of this law because more compounds are possible: as  $\text{CH}_4$ ,  $\text{C}_2\text{H}_6$ ,  $\text{C}_3\text{H}_8$ ,  $\text{C}_4\text{H}_{10}$ , etc.

**The Law of Combining Weights.**—When chemical substances are subjected to quantitative analysis, the relative weights of the elements entering into a compound are found. To use the mercury compounds as example, we find that the ratio of the weight of mercury to the weight of the chlorine in calomel is 199.8 to 35.18, while in corrosive sublimate the ratio is 199.8 to 70.36. We observe that the chlorine factor in the second compound is exactly twice that in the first compound. After analysis of large numbers of compounds the following law has been formulated: *Substances combine either in the ratio of their combining weights or in simple multiples of these numbers.*

#### SUMMARY OF CHAPTER V.

In any given chemical reaction, heat is either absorbed or liberated in an unchangeable ratio to the amounts of the substances reacting. The amount of heat absorbed when 1 gram of lime is made from hydrated lime is exactly the same as that given off when water is poured on 1 gram of lime.

If two substances react chemically and form a third,

they enter into combination in a constant proportion. When two elements combine in more than one proportion, the masses of the one which combine with a given mass of the other bear a simple rational relation to one another.

Substances combine either in the ratio of their combining weights or in simple multiples of these numbers.

## CHAPTER VI.

### ATOMIC WEIGHTS.

THE gas used to inflate balloons is hydrogen. This colorless, odorless gas is so light in weight that it causes the balloon to float in air. Hydrogen burns in oxygen to form an oxide—this oxide is water,  $H_2O$ . The gas chlorine of which we have spoken already as being able to combine with mercury also combines with hydrogen to form hydrogen chloride,  $HCl$ . The relative weights of the combining substances are: chlorine 35.18 to hydrogen 1; that is, 35.18 grams of chlorine unite with 1 gram of hydrogen to form 36.18 parts  $HCl$ . The substance iodine, with which we are all familiar also unites with hydrogen to form hydrogen iodide,  $HI$ , and in the ratio of 125.89 I to 1H. We remember that 199.8 parts mercury unites with 35.18 parts chlorine. By continuing our analyses beginning with any one element and finding the ratio of the combining weights of others, then taking these and working to still others we are able to obtain the combining weights of all elements in terms of one another. Now, how should we express them? The easiest way is to take the lightest element as our basis and let it equal 1. Since hydrogen is the lightest element we build our system on it stating that the combining weight is 1 and then chlorine will be 35.18; iodine 125.89; mercury 199.8, etc. In the hydrogen chloride combination we

find one in which the smallest quantity of hydrogen has entered. This 1 part of hydrogen must be at least one *atom*—it may be more but it cannot be less. By analyzing large numbers of compounds and by certain other methods we arrive at the conclusion that one atom of hydrogen combines with one atom of chlorine. Since the combining weights are respectively 1 (for H) and 35.18 (for Cl) we say that these numbers represent the *atomic weights* of these elements.

In chemical manipulations atomic weights play a very important part and if we can simplify the system of atomic weights to any extent we save the chemist that much. On the basis of hydrogen as 1 nearly all the atomic weights of the elements contain decimals. For example, oxygen when  $H=1$ , is 15.88. The accepted system is to let  $O=16.0$  to make it a round number, then H becomes 1.008 and arsenic which was 74.9 on the old basis now becomes 75.0; phosphorus, 30.96 becomes 31.0; mercury 199.8 becomes 200.0. This simplifies a great many calculations. It is usual then to state the atomic weight of oxygen as 16, though it makes no difference which system is used so far as the ultimate results are concerned, for all these numbers are relative. If we begin a calculation with either system we should of course use it throughout the immediate problem.

**Avogadro's Hypothesis.**—If one liter of hydrogen and one liter of chlorine are allowed to combine chemically there results not one liter of hydrogen chloride, but two liters of HCl.  $1L. H + 1L. Cl. = 2L. HCl$ . Now according to Avogadro's hypothesis "*In equal volumes*

*of all gases, at the same temperature and pressure, there is an equal number of molecules."* Therefore, in two liters of HCl there must be twice as many molecules as there are in one liter of H or one liter of Cl (or any other gas), and since each molecule must contain one atom of H and one atom of Cl it follows that each molecule of H and each molecule of Cl consists of two atoms. Then hydrogen as a gas exists not as H but as  $H_2$  and chlorine as  $Cl_2$ . This is in accordance with our previous statement that when atoms of elements have nothing else to combine with they sometimes combine with one another (see page 30).

**Molecular Weight.**—The sum of the atomic weights of the elements composing a molecule is called the molecular weight. For example, if a molecule of hydrogen is  $H_2$  and the atomic weight of H=1, then the molecular weight is 2; chlorine exists as  $Cl_2$ ; bromine as  $Br_2$ ; nitrogen as  $N_2$ , etc., therefore the molecular weights of these substances are twice their atomic weights. Mercury, however, exists as Hg, while in phosphorus there are three atoms in a molecule,  $P_3$ . Further, in substances composed of more than one element as HCl, the molecular weight is the sum of the atomic weight of H plus the atomic weight of Cl, or  $1.008 + 35.45 = 36.458$  = molecular weight of HCl.

Now for the reason why we learn about atomic weights and molecular weights. When chemical substances react with one another, they join atom to atom not gram to gram. Obviously, if hydrogen is the lightest substance then a gram would contain as many



more atoms as its weight is less than chlorine for instance—so that to obtain 36.458 grams HCl we do not mix 18.229 (one-half of 36.458) grams of H with the same quantity of chlorine but 1.008 gram H to 35.450 Cl. In using gaseous substances we can measure by volume but when we come to work with metals and salts we must weigh the materials.

#### SUMMARY OF CHAPTER VI.

The ratio of the units of combining power of elements is spoken of as *atomic weight*. Because hydrogen is the lightest gas known, it is accepted as a standard and its atomic weight placed, therefore, as 1. Other atomic weights are expressed in terms of hydrogen as 1. The atomic weight of O = 15.88 (*i. e.*, 15.88 times at. wt. of H). Recorded in terms of H many atomic weights contain decimals. To simplify calculations the value of H is placed at 1.008, then at. wt. of O = 16.0. Other examples are given in the text.

In equal volumes of all gases at the same temperature and pressure, there is an equal number of molecules (Avogadro's hypothesis).

Molecular weight is the sum of the weights of the atoms composing the molecule, *e. g.*, mol. wt. of HCl = at. wt. of H (1.008) + at. wt. of Cl (35.45) = 36.458. If the molecular weight of the compound having the empirical formula  $\text{CHO}_2$  is found by physical measurements to be approximately 90 and the sum of the atomic weights is 45 (C = 12, H = 1.008, O = 16) we learn that the real formula of this compound is  $(\text{CHO}_2)_2$ —or  $(\text{COOH})_2$ .

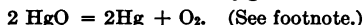
## CHAPTER VII.

### OXYGEN.

**Occurrence.**—It will be remembered that our experiments already cited in Chapter II proved that oxygen is one of the chief constituents of the air. It was also stated that the various metals occur in nature as oxides; for example, ordinary clay is the oxide of aluminum and silicon. Oxygen also exists in plant and animal bodies and 85 per cent. of the oceans of water is oxygen. Atmospheric air contains 23 per cent. pure oxygen. So abundant is the distribution of this element that it is estimated that one-half of the earth's crust is oxygen.

**Preparation in the Laboratory.**—Oxygen in air is mixed with nitrogen, a very inert (*i. e.*, will not combine easily) gas, and it is therefore difficult to separate them. One can decompose water by passing an electric current through it and hydrogen is given off at the *positive* electrode (where the current enters the water, so-called *anode*), and oxygen is given off as a gas at the pole where the current leaves, *negative* or *cathode*, and can be collected in an inverted tube.

The heating of an easily decomposable oxide as mercury oxide is a source of oxygen:



<sup>1</sup> The Latin for mercury is *hydrargyrum* and its abbreviation, Hg. One will wonder why this equation was not written thus:  $\text{HgO} = \text{Hg} + \text{O}$ . In the paragraph on Avogadro's hypothesis and mole-

Commercially oxygen is produced by heating equal parts of potassium chlorate and manganese dioxide. The gas given off is purified by washing in different solutions and compressed into tanks.

**Uses.**—Tubes or tanks of oxygen are found in every hospital for emergency use in cases of asphyxiation and where there is small working lung space as in pneumonia or diminished breathing capacity. Oxygen is necessary in nitrous oxide anesthesia.

In the arts oxygen is used with illuminating or acetylene gas to produce a flame of intense heating power. With such a flame one can cut through steel plate with the greatest ease.

**Properties of Oxygen.**—Oxygen is a transparent gas, possessing neither color nor odor, and weighs one and one-tenth times as much as air. By diminishing the temperature to an extreme degree of coldness (below  $-119^{\circ}\text{C.}$ ) and exerting very high pressure, oxygen may be liquefied, though with such difficulty that for a long time it was considered impossible.

The most familiar chemical phenomenon is combustion which has already been referred to as *oxidation*. It makes no difference whether the reactions take place quickly as the burning of wood, or proceeds slowly, as the tarnishing of silver or the rusting of iron, the chemical process is the same, namely, the chemical

cules it was stated that certain atoms combine with one another when they have nothing else to combine with. This is true of O, that is, O exists as  $\text{O}_2$  though mercury exists as Hg. Then, so far as Hg is concerned, this last equation is possible, but if O must combine with something, it would return to the Hg so soon as it is released unless there is another O present to form  $\text{O}_2$ . We therefore presume that these processes go in pairs, that is,  $2\text{HgO} = 2\text{Hg} + \text{O}_2$ .

union with oxygen. In pure oxygen these reactions take place more easily than in air which is oxygen diluted with an inert gas (nitrogen). For example, a piece of phosphorus exposed to the air will take fire spontaneously and burn quietly, but if placed in pure oxygen, oxidation takes place immediately and with explosive violence. The reason for this is found in the fact that chemical reactions proceed more rapidly as we increase the temperature. When the gas is pure oxygen, all the heat of the chemical reaction goes to raise the temperature of the reacting substances, whereas in a mixture like air a greater part of the heat is absorbed by a non-reacting substance (nitrogen).

**Oxides.**—When an element burns or is slowly acted upon by oxygen an *oxide* is formed. Certain elements are capable of combining with oxygen in different proportions forming totally different compounds. By burning carbon in various amounts of air, we are able to produce a compound containing 1 part of oxygen for every carbon atom and also to prepare an oxide containing two parts of oxygen for every carbon atom. If there is an excess of carbon and very little air the former oxide (CO) called *carbon monoxide*, results, while with abundant air supply *carbon dioxide* (CO<sub>2</sub>) is produced. Carbon monoxide is a deadly poisonous gas while carbon dioxide is relatively harmless.

Thus the oxides are named first according to the number of atoms of oxygen present as *monoxide*, *dioxide*, *trioxide*, *pentoxide*, etc. Where the number of atoms of the element in combination with the oxygen may also be different we use “ous” and “ic” to indicate oxides

poorer or richer in oxygen.  $\text{FeO}$  is called *ferrous* oxide, while  $\text{Fe}_2\text{O}_3$  is *ferric* oxide—an *-ous* oxide contains less oxygen while *-ic* oxides contain more. We shall find that this same terminology is applied to other compounds like the chlorides, iodides, sulphides, etc.

**Ozone.**—The peculiar pungent odor noticeable in the neighborhood of electrical dynamos and *x*-ray apparatus is due in part to the presence of ozone in the air. It is best produced by passing a silent discharge of electricity through pure oxygen. Ozone is a very active oxidizing agent and for this reason is a very efficient deodorizer and disinfectant. It is used to rid municipal water supplies of bacteria by allowing the water to fall from a tower while ozone is bubbled through it.

A few hospitals in America use ozone to sterilize their operating room supplies. Metallic silver when placed in an atmosphere of ozone at ordinary temperature will be covered with a layer of brown oxide; while oxygen under similar conditions will not bring about such an oxidation. By heating to  $300^\circ \text{C}$ . ozone is changed to oxygen—which in turn may be retransformed to ozone by an electric current. On account of this and other facts it is believed that ozone is oxygen plus an extra amount of energy and instead of existing as  $\text{O}_2$  (oxygen gas) is  $\text{O}_3$ .

#### SUMMARY OF CHAPTER VII.

Oxygen is a transparent gas possessing neither color nor odor. It is slightly heavier than air and can be liquefied. Oxygen is necessary for life. It is very

widely distributed in nature (about half the crust of the earth is said to be oxygen bound to metals and other elements forming oxides). Combustion is an oxidizing process, that is, the union of oxygen with any substance. Oxidation may proceed rapidly with the product of light and heat (fire) or very slowly in water solution (life processes).

Oxygen may be prepared by decomposing water with an electric current or by heating mercury oxide.

Ozone is a pungent gas product by electric sparks in oxygen. There are reasons for the belief that ozone is oxygen plus energy and that the molecule of oxygen gas contains two atoms ( $O_2$ ) while the ozone molecule contains three atoms ( $O_3$ ). Ozone is used to sterilize water supplies and as an oxidizing agent in the laboratory. Ozone is also an efficient deodorizer.

## CHAPTER VIII.

### HYDROGEN, $H_2$ .

(Mol. wt. = 2.016; At. wt. = 1.008.)

**Occurrence.**—Very small quantities of hydrogen occur in the atmosphere of the earth but by spectrum analysis it has been discovered to be very widely distributed in the stars, the sun, and is found in nebulous masses. Deposits of free hydrogen have been found in great salt deposits and in oil wells. Eleven per cent. of water is hydrogen, so that in actual amounts it ranks next to oxygen in the earth's composition.

**Preparation in the Laboratory.**—Just as oxygen is given off at the *positive* pole when an electric current is passed through water, hydrogen the only other component is given off at the *negative*<sup>1</sup> electrode and may be collected in the closed end of a tube.

The practical method of producing hydrogen is to

<sup>1</sup> Since *positively* charged substances attract *negatively* charged substances and *vice versa* we assume that oxygen is a *negatively* charged substance because it is attracted to the *positive* pole and that hydrogen is *positively* charged because it is attracted to the *negative* pole. We shall later learn that the metals (iron, copper, silver, nickel, sodium potassium, etc.) are like hydrogen in being classed as *positive* elements, while chlorine, bromine, iodine, sulphur, fluorine can be classed as *negative* as a general rule and that one of the *positive* elements will easily combine with a *negative* element. Some elements as arsenic and phosphorus combine easily with either oxygen or hydrogen but to attempt to explain this would take us into physical chemistry.

allow hydrochloric acid (the acid which is normally present in the stomach) to act on a metal like zinc:



**Uses.**—Hydrogen is compressed into tanks. It burns with an intense non-luminous flame and hence is used in welding and heating processes. With a hydrogen-oxygen blow-pipe cast iron can be welded. Balloons are inflated with hydrogen because it is lighter than air. In the laboratory hydrogen is used as a reducing agent, that is, used to combine with the oxygen held in combination by some other substance and thus de-oxidize. In this case it is, as a rule, generated in contact with the substance to be reduced for it is more active just at the point of liberation when it is said to be *nascent* (born).<sup>1</sup> In the bacteriological laboratory hydrogen is used to displace the oxygen of the air when it is desirable to cultivate organisms which will not grow in the presence of oxygen.

**Properties of Hydrogen.**—Hydrogen is a transparent gas possessing neither color nor odor and is the lightest of all known substances. It can be liquefied with difficulty by diminishing the temperature to  $-200^\circ$  under a pressure of three hundred atmospheres and allowing it to expand into a pressure of about fifty atmospheres when the heat absorbed during the process of expansion will still further cool it until it liquefies. If liquid hydrogen is poured into a test-tube liquid air will flow from the outside and finally the tube will be covered with frozen air.

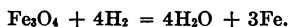
<sup>1</sup> The H atoms are liberated and seek to combine with oxygen before combining with another H to form  $\text{H}_2$ .



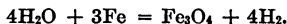
Hydrogen combines with chlorine, iodine, bromine to form acids, in fact hydrogen is present in, and is a necessary constituent of, all acids. In combination with nitrogen, hydrogen forms ammonia ( $\text{NH}_3$ ) and with carbon forms a long series of so-called *organic* compounds. Hydrogen combines readily with sulphur to form the bad-smelling gas hydrogen sulphide  $\text{H}_2\text{S}$ , and with palladium and sodium, etc., to form *hydrides*,  $\text{Pd}_2\text{H}$ ,  $\text{NaH}$ .

Hydrogen seems to possess a special affinity for oxygen. If these two gases are mixed in the ratio of two parts of hydrogen to one of oxygen they form a highly explosive mixture which can be easily set off by a flame. This fact should be remembered when working with hydrogen gas.

When hydrogen gas is passed over an oxide of a metal such as iron at high temperatures, the hydrogen combines with the oxygen to form water as in the following equation:



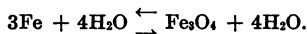
If water vapor is passed over heated iron the reverse is true:



How then are we going to know what will happen? It has been found that if hydrogen is present in excess metallic iron will result, but if water vapor is present in excess the oxide of iron is formed.

This illustrates the effect of *quantity* (mass) on chemical reactions and the reaction here given is said to be *reversible* in that it may proceed either way

according to the quantity of materials present. To indicate the reversibility of chemical reaction the equation may be written thus:



Most chemical reactions are *reversible*.

### SUMMARY OF CHAPTER VIII.

Hydrogen is an element. It is a transparent, colorless, odorless gas and the lightest of all known substances. It can be liquefied. Hydrogen combines with chlorine, iodine, bromine, etc., to form acids; it unites with metals to form hydrides, and with carbon and oxygen to form a long list of organic compounds. It unites with nitrogen to form ammonia ( $\text{NH}_3$ ) and with oxygen to form water ( $\text{H}_2\text{O}$ ).

Just as the addition of oxygen to a substance is called oxidation the addition of hydrogen is called reduction. When hydrogen is added to a substance containing oxygen, the hydrogen and oxygen unite to form water and split off from the molecule of the substance—so that reduction may mean the abstraction of oxygen by hydrogen. If no oxygen is present reduction means simply the addition of hydrogen to the molecule.

Hydrogen occurs in small quantities in the atmosphere and free in the earth's crust. It is present in large amounts in combination: water is 11 per cent. hydrogen (by weight).

Hydrogen results from the decomposition of water by an electric current. The volume of the hydrogen given off at the negative pole is twice the volume of the

oxygen given off at the positive pole. Hydrogen is said to be positive because it is attracted to the negative pole.

Hydrogen is also prepared by the action of HCl on zinc.

Hydrogen is used in balloons (on account of its lightness) and in chemistry for reduction. It is used to replace oxygen in certain bacteriological methods.

## CHAPTER IX.

### WATER, $H_2O$ .

(Mol. wt. = 18.)

WHEN pure hydrogen or any compound containing hydrogen is burned in oxygen or air, water results. If a cold object is held above a flame water condenses on it. This is true if the flame consists of burning gas which has been deprived of all water vapor, so it follows that water is one of the end-products of combustion.

**Occurrence.**—By far the commonest chemical compound is water. Widely distributed as it is over and through the earth's crust, the chief constituent of animal and vegetable matter, it is so common and so familiar that we accept it without inquiry as to its composition. Not until we begin the study of chemistry do we wonder what its makeup is, and since chemistry can tell us about such common things this science no longer seems artificial and set apart from every-day things. From now on we shall look upon water, salt, wood, rocks and coal in a new light—that is, from the stand-point of their chemical composition.

Three-fourths of the earth's surface is covered with water, 70 per cent. of our bodies is water, and stones which seem to be dry contain an astonishingly large amount of it. Without it the earth would be dead.

**Uses.**—The great solvent for chemical substances is water. Without it many chemical substances will not react; for example, an explosive mixture of hydrogen and oxygen will not explode if both gases are perfectly dry; hydrogen and chlorine will remain forever uncombined if mixed in an absolutely dry state. All chemical processes of living matter, plant and animal, take place in the presence of water. Digestion, assimilation, oxidation and elimination require water and by the aid of it heat regulation of the body is possible. The decompositions of plants and animals by which the elements are liberated to be reformed into living matter require moisture, since these processes are carried out by bacteria and their ferments which cannot work in the absence of water. Water is the great cleanser and heat regulator of the earth.

**Properties.**—The physical properties of water are so well known that it is not necessary to review them here. It is wise, however, to discuss certain facts concerning such a common substance in order that other less common substances and their forms may be compared with it. On account of the fact that water is so easily obtained in a pure condition, it is used as a standard of comparison.

The scientific standard of distance is the *meter* (about 39 inches).

One-hundredth of a meter is called a centimeter. The volume of a perfect cube which measures exactly 1 cm. on every side is 1 c.c.; in other words, the amount of liquid which a cubic container measuring inside 1 cm. each way is 1 c.c. From this measurement

of distance we obtain standards for weight measure by using pure water. At 15° centigrade 1 c.c. of water weighs exactly 1 gram, which is the standard for weight measure. (See Weights and Measures, p. 24.)

**Specific Gravity.**—(Sp. gr.) Specific Gravity means *particular weight*. It is a well known fact that a gallon of molasses weighs more than a gallon of water, that is, volume for volume molasses is heavier than water. If 1 c.c. of molasses weighs 1.5 gram and 1 c.c. pure water weighs 1 gram then obviously volume for volume molasses is one and one-half times as heavy as water, or accepting water as the standard of comparison with *specific gravity* as 1.000 then the *specific gravity* of molasses is 1.500.

Molasses is a solution of sugars in water with some caramel and extractions which add to the color and flavor. The more sugars dissolved in a given volume of water the greater the specific gravity. When we come to study the chemistry of the urine, we shall see that the determination of specific gravity is very important for the reason that we are able to ascertain by a very quick and simple process the amount of solids in the urine. The salts and nitrogen-containing bodies dissolved in it make its specific gravity vary from 1.020 to 1.030. High specific gravity leads us to suspect the presence of sugar in urine.

**Hydrometer.**—The instrument for quick determination of the specific gravity of liquids consists of a glass bulb filled with air and weighted with mercury so that the 1.000 mark on the graduated stem is exactly at the surface when the instrument floats in pure water.

It is called the *hydrometer* (water-measure). See chapter on Uranalysis.

**Lighter Liquids.**—Ether floats on water and is therefore lighter than water. 1 c.c. weights only 0.717 gram: its specific gravity is 0.717. Pure alcohol has a specific gravity of 0.797, but when water is added its specific gravity is proportionally increased. By determining the specific gravity of a sample of alcohol one can ascertain how much water has been added. Kerosene and gasoline are obtained from the same source, both are lighter than water, but gasoline has the lower specific gravity. These two liquids are sold on the basis of their specific gravity. If the specific gravity of kerosene is low, it indicates that it contains too much gasoline and is therefore dangerous for use in lamps, and if the specific gravity of gasoline is too high it is not fit for motors. We see that the determination of specific gravity has a very practical bearing, since we may employ it in determining the relative purity of certain liquids.

**Thermometer.**—The freezing- and the boiling-point of pure water are the constants upon which the standard chemical thermometer is made. As the name (Centigrade, 100°) indicates the difference between the boiling-point and the freezing-point is divided into one hundred parts and each part called a degree. Then water boils at 100° C. and freezes at 0° C. The only other kind of thermometer with which nurses must be familiar is the Fahrenheit, named after its inventor. Water boils at 212° F. and freezes at 32° F. The difference between the two is 180° F., and it follows that, since

100° Centigrade covers the same range as 180° F., 1° C. = 1.8° F. If we wish to convert Centigrade to Fahrenheit we multiply by 1.8 and add 32 (because in Centigrade we begin at freezing-point), which equals 0.37° C., means 37° C. above freezing-point, whereas if 37° C. equals in range 66.4° F. we must take it to mean 66.4° above the freezing-point of water. Since 32° F. is freezing-point of water, 66.4° above this point is  $66.4 + 32.0 = 98.4$ . Therefore 37° C. = 98.4° F.<sup>1</sup>

To convert Fahrenheit reading to Centigrade we subtract 32 and divide by 1.8.<sup>2</sup>

**Boiling-point.**—It has been stated that the boiling-point of pure water is 100° C. This is true at sea level, but if we go up a mountain we find that the boiling-point gradually becomes less and less. On the top of Mont Blanc the boiling-point is sometimes as low as 84°. The reason for this lowering is the decrease in atmospheric pressure. The same result may be accomplished in the laboratory by reducing the atmospheric pressure by means of a vacuum pump. Knowing the boiling-point of water at any place we are able to estimate roughly the height above sea level.

If, on the other hand, we increase the pressure on the surface of the water the boiling-point is correspondingly raised. This is what happens in a steam boiler or steam sterilizer. In the latter instead of a

<sup>1</sup> 37° C. or 98.4° F. is the normal temperature of the human body.

<sup>2</sup> Since the Centigrade thermometer is being used more and more in clinical work, and since the confusion of the two systems in carrying out orders might at times be dangerous, it is suggested that the pupil nurses convert several readings, one into the other. Room temperature 68° F. (20° C.), normal body temperature and bath temperatures are good points to fix in their minds.



temperature of 212° F., we reach 225° F.; but the pressure is also increased to about 15 pounds per square inch. On account of being able to increase the temperature above the sea-level boiling-point we are enabled to kill many bacteria not killed by boiling.

#### SUMMARY OF CHAPTER IX.

Water is composed of two volumes of hydrogen and one volume of oxygen ( $H_2O$ ). By weight 11 per cent. of water is hydrogen and 89 per cent. oxygen.

When hydrogen or any compound containing hydrogen is burned in air (or oxygen) water is one of the products of combustion.

Water is necessary for life and for many chemical reactions. It is the great chemical solvent.

Water is very widely distributed in nature. It can be easily obtained in a pure state, and is therefore used as a standard of comparison for various physical characteristics. The relative weight of a given volume of any substance compared with an equal volume of water is spoken of as the specific gravity of a substance. The specific gravity of water is taken as 1.0.

The standard unit of volume is the cubic centimeter. 1 c.c. weighs 1 gram at 15° C.

The instrument for measuring the specific gravity of liquids is the hydrometer.

The thermometer registers the *degree* of heat. There are several kinds; especial mention is made of the Fahrenheit and the Centigrade. Water freezes at 0° C. or 32° F. Water boils at 100° C. or 212° F. To convert Centigrade reading to Fahrenheit multiply

by  $\frac{9}{5}$  (or 1.8) and add 32. To convert Fahrenheit readings to Centigrade subtract 32 and either multiply by  $\frac{5}{9}$  (or 0.555) or divide by 1.8.

Boiling-points of liquids decrease with a decrease in pressure and *vice versa*. Water boils below 100° C. (212° F.) on top of mountains.

## CHAPTER X.

### HEAT.

**Heat Absorbed in Evaporation.**—The thermometer measures the *degree* of heat not the amount of heat. Obviously it requires a larger amount of heat to raise a gallon of water  $1^{\circ}$  than it takes to raise the temperature of a quart of water  $1^{\circ}$ . Just as we have liquid measures and standards of weight we accept an easy and simple standard for measuring heat. The standard unit is the Calorie<sup>1</sup> (large Calorie) which is the heat required to raise the temperature of 1000 grams of pure water  $1^{\circ}$  C. One thousand grams (1 kilogram) of water equals 1000 c.c. or 1 liter and is approximately one quart, so that a Calorie is approximately the amount of heat necessary to raise the temperature of a quart of water  $1^{\circ}$  C. or  $1.8^{\circ}$  F.

To raise the temperature of one quart of water from room temperature ( $20^{\circ}$  C.) to the boiling-point ( $100^{\circ}$  C.) about 80 Calories is needed. Now in order to convert this amount of water at  $100^{\circ}$  C. into steam at  $100^{\circ}$  C. a great amount of heat is necessary, viz., 540 Calories. This heat necessary to change a substance from a liquid to a gaseous state is known as the *heat of vaporization*. The heat of vaporization is absorbed by the steam and given out again on condensation. The heat

<sup>1</sup> The small Calorie is the heat required to raise the temperature of 1 gram water  $1^{\circ}$  C.

of vaporization interests us on account of the fact that nature makes use of it in the cooling of the human body; sweat is secreted upon the skin by the sweat glands and the large amount of heat absorbed during vaporization lowers the temperature of the body.

The definition of the term Calorie should be remembered, for use will be made of it later in the discussion of food values.

**The Freezing of Water.**—The well known principle: heat expands, cold contracts, holds good for water between certain limits. On being heated water expands slightly and when the same mass occupies more volume the density or specific gravity becomes less. For this reason specific gravity determinations and accurate measurements of water must be made at an accepted temperature—say  $15^{\circ}\text{C}.$ , in order to be comparable.<sup>1</sup>

The differences in density thus produced cause up and down currents in bodies of water.<sup>2</sup>

Water becomes denser on cooling until the temperature of  $4^{\circ}\text{C}.$  is reached and between  $4^{\circ}$  and  $0^{\circ}$  it expands, so that when it freezes the ice produced is lighter than water and will float. If water continued

<sup>1</sup> Sometimes determinations of specific gravity are made at  $25^{\circ}\text{C}.$  based on the density of water at  $25^{\circ}\text{C}.$ , then the result of such a determination of specific gravity of, let us say, a sample of urine would be expressed thus:  $\text{Sp. gr. } 25^{\circ}/25^{\circ} = 1.0236$ . This means that the sample of urine at  $25^{\circ}\text{C}.$  is 1.0236 times as heavy as water at  $25^{\circ}\text{C}.$

<sup>2</sup> In the spring and autumn one often notices a peculiar fishy, violet, or aromatic odor in tap water. This is due to the stirring up of algæ by the spring or autumn *turn-over*. These turn-overs result from the difference in density of the top and bottom layers of bodies of water—the top cooling, thus becoming more dense settles while the bottom layer comes up.

to contract to  $0^{\circ}$  ice would form and sink to the bottom. Streams and lakes would freeze throughout.

**Heat Absorbed in Melting.**—Ice at  $0^{\circ}$  cannot change to water at  $0^{\circ}$  without absorbing a relatively large amount of heat. The amount of heat absorbed by one kilogram of ice (2.2 lbs.) in changing to water is about 80 Calories, or, the amount of heat abstracted from a liter (about 1 quart) of water in being cooled from  $80^{\circ}$  C. to zero. Ordinary room temperature is from  $20^{\circ}$  to  $25^{\circ}$  C., and refrigerator temperature varies from  $4^{\circ}$  C. to  $14^{\circ}$  C., a difference of, let us say,  $10^{\circ}$  C. Then one kilogram of ice in melting would absorb enough heat to cool over 8 quarts of water from room temperature to refrigerator temperature.

**Freezing Mixture.**—When ice is placed in water ice melts and cools the water until there is an equilibrium established between the solid and liquid forms of water at  $0^{\circ}$  C. When more heat is absorbed more ice goes into the liquid form and the temperature is lowered. If one adds salt to ice or the water-ice mixture it has a tendency to make more ice melt, and when ice melts it absorbs heat; therefore the water-ice-salt mixture becomes colder. If an excess of ice and an excess of salt be present a constant temperature of  $-18^{\circ}$  C. ( $18^{\circ}$  C. below freezing point) is maintained. If the temperature is higher than  $-18^{\circ}$  C. more salt goes into solution and melts more ice which absorbs heat until the constant point ( $-18^{\circ}$ ) is reached. If the mixture cools lower than  $-18^{\circ}$  salt crystallizes out of solution and allows ice to form until the temperature is raised to a constant point.

The reason for these facts cannot be explained without going into physical chemistry but the application can be readily seen. This is the method used in making ice-cream or for cooling anything quickly. The essential thing in making such a freezing mixture is to always have present an excess of salt and an excess of ice. It should be remembered that  $-18^{\circ}$  C. is intense cold and that such a mixture should not therefore be used in an ice-bag applied to a patient.

#### SUMMARY OF CHAPTER X.

The thermometer registers the degree of heat. The quantity of heat is expressed in Calories. A small Calorie is the amount of heat necessary to raise the temperature of one gram of water  $1^{\circ}$  C. The large Calorie (C) equals 1000 small Calories. The large Calorie is generally used.

The heat necessary to change a substance from a liquid to a gaseous state (water at  $100^{\circ}$  C. into steam at  $100^{\circ}$  C.) is known as the *heat of vaporization*.

Heat of vaporization is exactly equal to *heat of condensation*.

One kilogram of ice at  $0^{\circ}$  C. absorbs 80 Calories in changing to water at  $0^{\circ}$  C.

Heat of freezing is exactly equal to heat of melting.

A mixture of ice-salt-water will maintain a temperature of  $-18^{\circ}$  C. so long as ice and salt are present in excess.

## CHAPTER XI.

### SOLUTIONS AND PURIFICATION OF SUBSTANCES.

#### SOLUTIONS.

**WATER** is the most important solvent. According to Jones' Inorganic Chemistry over three-fourths of the chemical reactions with which we are familiar take place in water solution; certainly the reactions which take place in nature proceed in aqueous solutions.

Nearly all substances—solids, liquids and gases—are soluble to a greater or less degree in water. Many things are regarded in a practical sense as being insoluble in water, which are so slightly soluble as to be negligible, nevertheless possess a definite degree of solubility. We are familiar with the solution of solid things in a liquid but, perhaps, we have not thought of the insolubility of liquids. Alcohol and water are miscible in all proportions, that is, the one has unlimited solubility for the other; but chloroform and water will not mix in all proportions. Pour some chloroform into water and it sinks to the bottom. Shake them and the chloroform is emulsified but soon collects again at the bottom of the vessels but in a smaller amount. The water if poured off will retain a definite amount of chloroform and has acquired a sweetish taste and an odor characteristic of chloroform. This is the way

chloroform liniment is prepared. By determining the specific gravity of the solution we find that it is greater than 1, and knowing the specific gravity of chloroform we calculate using the exact figures found, the exact amount of chloroform dissolved.

Gases are soluble in water according to the pressure of that particular gas on the surface of the liquid. The presence of other gases creating an enormous pressure would have no effect on this.

Heat some water from the tap in a glass beaker and observe the bubbles coming off long before the water reaches the boiling point. This is the dissolved air being driven off. Taste some water which has been recently boiled and observe its flatness. Agitate it with air and it regains its normal taste.

**Effect of Temperature on Solubility.**—The general effect of raising the temperature of the solvent is to increase its power to dissolve. Therefore if we wish to hasten the preparation of a solution or to make a stronger solution we heat the solvent. As has been stated in the first chapter some salts take up heat when they are dissolved; that is, the solution gets colder as more salt is dissolved and consequently solution is slow. This is true of magnesium sulphate (Epsom salts), so that when making a solution of this chemical it will shorten the process materially if warm water and heat are used.

Gases are *less* soluble as the temperature is raised, therefore solutions of gases are kept cold. In effervescing magnesium citrate the carbon dioxide is in solution in the bottle on account of the pressure.



When the bottle is opened and pressure released the gas comes out of solution causing the foam. If the citrate solution is warm practically all the gas is lost, while in cold solutions more gas is retained.

### PURIFICATION OF SUBSTANCES.

**Crystallization.**—The difference in degree of solubility is made use of in the separation and purification of salts. An earth containing mixtures of salts is leached, the solutions filtered and concentrated to the point of crystallization. When a crystal forms its tendency is to come out pure so that by repeated crystallization salts may usually be freed of their impurities.

Water is used extensively as the solvent in the purification of substances, but when the salts have about the same degree of solubility in water, other solvents are made use of.

Water itself can be purified by successive crystallization. In the changing of water to its solid form (freezing), ice separates out purer than the mother liquor each time. Many substances which either crystallize with difficulty at very low temperatures or not at all are purified by distillation.

**Distillation.**—Distillation is the most widely used and easiest method of purifying water<sup>1</sup> for laboratory

<sup>1</sup> One distillation is not sufficient to obtain absolutely pure water. The first and last portions are thrown away since they generally contain soluble gases and ammonia. Water distilled in glass contains a trace of alkali (lye) and that from black tin stills contains a trace of tin. Certain bacteria grow in distilled water to a slight extent and even if killed by heat they produce ill effects when injected into the body. Therefore for injection purposes (salvarsan, saline, etc.), water should be double distilled and freshly prepared.

purposes, etc. Various oils and alcohol are freed of impurities by this method. In this process the differences in boiling-point are made use of: distillation is allowed to proceed at a given temperature until all the vapor that comes off at that temperature has been recovered and the temperature is then raised to obtain the next fraction. Such a process is termed *fractional distillation*. Curiously enough, mercury is separated from substances with which it occurs in nature by being distilled, and the same process is used to further purify this metal in the laboratory.

#### SUMMARY OF CHAPTER XI.

Nearly all substances are soluble to a greater or less extent in water. A large proportion of chemical reactions take place in water solutions.

As a rule, heating increases the solvent power of water for liquids and solids and decreases its solvent power for gases.

Other things being equal the amount of gas dissolved in a liquid varies as the pressure of that particular gas on the surface.

The difference in solubilities of two substances in any given solvent make it possible to separate the substances by recrystallization. Substances may be freed of impurities by recrystallization if the impurities possess a different degree of solubility.

Liquids having different boiling-points may be separated by distillation. Distillation offers another means of purifying substances.

## CHAPTER XII.

### NATURAL WATERS—CHEMICAL ACTION OF WATER.

**Sea Water.**—Owing to its power of dissolving substances, water as it occurs in nature is almost always impure. Sea water contains large amounts of salts which the waters falling on the surface of the earth and seeping through the crust have leached out and brought with them. The air takes up the water in vapor form and leaves the salts. The Dead Sea which has no outlet to the ocean contains a very large quantity of salts (about 25 per cent.), so much that the specific gravity is very high and it is impossible for a person to sink in it.

**Lakes, Rivers, Springs.**—The amount of impurities in the waters of springs, rivers and lakes depends primarily upon the character of the region through which the water flowed. If a spring is furnished by waters flowing through sandstones and quartz formations it will contain very little impurities, while springs from salt deposits will, of course, contain varying amounts of minerals. Springs coming from very deep sources where the temperature is high, will, when other things are equal, contain more mineral on account of the increase in solvent power due to increase in temperature. Surface waters are liable to contamination by all sorts of impurities chief among them are the excreta of man and animals. Inland

waters containing relatively large amounts of common salt without other minerals in proportion are looked upon by the sanitarian as possibly contaminated by man or animals. Other tests are necessary to confirm this suspicion.

**Hard and Soft Waters.**—Waters slowly passing through swamps take up organic acids which increase the solvent action. Such waters passing over limestone will take up relatively large quantities of lime and become what are known as *hard* waters.

Hard waters coagulate soap solutions, that is, when soap is used in them they become turbid and if very hard will precipitate the soap in flakes or curds. If the hardness is due to limestone (calcium carbonate) and magnesium carbonate, boiling will drive off the carbonic acid and allow the salts to precipitate and the water is *soft*. This is called *temporary* hardness. If the water contains calcium sulphate or iron sulphate (copperas) it is said to be *permanently* hard. It is plain that the use of hard waters for washing purposes is not economical because so much soap is needed to produce any effect. The remedies generally used to soften waters are: boiling or the addition of borax (sodium baborate).

Waters containing salts in any appreciable amounts should not be used for boilers or for steam sterilizers (autoclaves) because of the residue they leave to form a cake. Waters containing organic acids corrode the sterilizers and even hasten the rusting of instruments. A small amount of sodium carbonate (washing soda) is added to the water in which instruments are

boiled in order to neutralize any acids that may be in the water and thus prevent corroding and rusting.

**Rain Water.**—Rain water is soft and one would think it ought to be pure. Rain is the cleanser of the atmosphere and brings down besides the dust around which rain drops form certain gases like ammonia.

**Drinking Water (Potable Water).**—The fact that a water is clear and cold and contains very small amounts of solid matter does not necessarily mean that it should be drunk. One cannot judge the potability of water entirely by its taste, odor, or appearance. Apparently the best water may contain disease-producing germs like typhoid, cholera, etc., so that a water to be safe must conform to certain bacteriological, as well as chemical and physical standards. Fortunately filtration will also remove a large percentage of the bacteria. Mechanical filtration, after the addition of clarifiers, is often used, but the safest filter is slow sand filtration. The latter is a biological as well as a physical and chemical process. Each grain of sand is coated with a gelatinous membrane due to the growth of organisms and the bacteria coming through impinge upon them and are held. An efficient sand filter removes 97 to 99 per cent. of the germs of the water passing through them.

Porcelain filters through which water is forced under pressure are made for home use as well as for the laboratory. Of this type the Berkefeld and the Pasteur are commonly used. They are made in various degrees of fineness and standardized according to the amount of water which will pass through in a given

time under a given pressure. No filter can be said to be safe until it is tested bacteriologically, and then during an epidemic of a water-borne disease even the filtered water should be boiled before being consumed.

#### SUMMARY OF CHAPTER XII.

Natural waters are almost always impure. The amount of salts and other impurities in spring, river or lake waters depends upon the region through which the water flowed. Sea water and mineral springs contain large amounts of salts. Hot springs and deep well waters are liable to contain large amounts of salts.

Water containing large amounts of limestone and magnesia is said to be *hard*. The lime and magnesia form insoluble salts with soaps and no lather is formed until all the salts are precipitated. The use of hard waters for bathing purposes is uneconomical. Waters may be softened by boiling or the addition of borax. Waters containing calcium and magnesium carbonate are said to be in a state of *temporary* hardness because this property can be removed by boiling. *Permanently* hard waters contain magnesium, calcium or iron sulphate. Rain water and distilled water are soft. Waters from swamps passing over limestone are generally hard.

Waters containing salts should not be used in autoclaves or boilers on account of their scale-forming properties.

The potability of waters cannot be determined solely by appearance and taste. Polluted waters may appear highly desirable for drinking purposes.

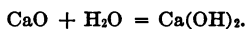
## CHAPTER XIII.

### COMPOSITION OF WATER.

THE composition of a substance is determined by observations from two stand-points—analysis and synthesis. By the analysis of water accomplished by passing an electric current through it we find that two elements are obtained. Hydrogen is given off at one pole (negative) in twice the amount in which oxygen is given off at the other (positive) pole. This shows us that at least water is composed of two parts of hydrogen to one of oxygen or  $\text{H}_2\text{O}$ . But there may be other substances liberated which do not come off as a gas, so we test the water remaining undecomposed in the flask but find no other substances. Then we make a determination of the molecular weight of water vapor by certain physical chemical means and find it to agree with the formula  $\text{H}_2\text{O}$ . Now, if our work has been correct we ought to be able to *synthesize* water from these two elements in these proportions. So two parts of pure hydrogen gas and one part of pure oxygen gas are placed in a vessel and exploded by means of an electric spark. Water is formed and all of both gases disappear. We conclude that water is composed of two parts of hydrogen to one part of oxygen.

This reaction gives off a great amount of heat and

the compound formed, water, is very stable. It is very difficult to decompose water by heat. In fact water is a very stable compound though it may combine chemically with certain substances with apparent ease it reacts as a whole molecule and is not broken up. It combines with oxides to form the hydroxides such as was indicated in the first chapter. Lime which is the oxide of calcium (CaO) is converted into lime hydroxide or calcium hydroxide (Ca(OH)<sub>2</sub>) by water thus:



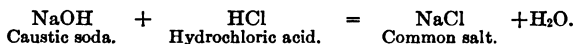
Water also combines with the dioxide (SO<sub>2</sub>) or dioxide (SO<sub>3</sub>) of sulphur to form acids:



and



Water is probably the most important chemical compound known. It is a by-product in great many reactions in which hydrogen combines with an OH group. For instance, in the union of an acid and a base about which we shall learn later, water as well as a salt are formed:



In this case the H of the acid combined with the OH of the base. This OH group is very important in chemistry and should be remembered for use later. It is called the *hydroxyl* group.

**Ionization.**—Substances in solution in water do not exist there as closely bound elements but are slightly



pulled apart; as:  $\text{H—Cl}$  or  $\text{Na—OH}$ . The compound is said to be *ionized* or in loose formation (dissociated). Each component is an *ion* and termed positive or negative. The extent of dissociation depends upon the dilution, that is, in weak solutions, we find more dissociation or ionization than in strong solutions.

When two or more salts are in the same solution there is a constant interchange of ions. Suppose, for example, that saltpetre (potassium nitrate  $\text{KNO}_3$ ) is in solution with common salt (sodium chloride,  $\text{NaCl}$ ). The former would be dissociated into  $\text{K—NO}_3$  and the latter into  $\text{Na—Cl}$ . Immediately there would be an interchange with the loose formation of potassium chloride,  $\text{K—Cl}$  and of sodium nitrate  $\text{Na—NO}_3$ . We may imagine a sort of a Virginia reel formation, the girls representing the negative ions and the boys the positive ions. The positive ions would swing his partner then swing some other negative ion, back to his particular negative ion and then to some other and so on so long as all combinations remained in the game. Let us introduce another chemical like lunar caustic (silver nitrate  $\text{AgNO}_3$ ). When the positive ion  $\text{Ag}$  is in loose combination with any  $\text{NO}_3$  ion the interchange proceeds as usual but when an  $\text{Ag}$  ion comes in contact with a  $\text{Cl}$  ion they form a compound which is insoluble and precipitate from solution. In terms of our illustration this particular combination of  $\text{Ag}$  and  $\text{Cl}$  leave the game immediately and this precipitation proceeds until either all the  $\text{Ag}$  or the  $\text{Cl}$  is used up. Either may be in excess and use up all the other and precipitation ceases. This fact is made

use of in the determination of the amounts of either material in solution. If we have a chloride in solution and want to know how much chlorine there is present, we add a soluble salt of silver in excess and the chlorine is precipitated as silver chloride. This is filtered off, washed, and weighed. If we have silver in solution—some soluble chloride is added, and the precipitate treated as above.

The question arises: How are we to know which combination will be insoluble and will precipitate? The answer is: We cannot predict what will happen except by experience. We learn accidentally or by trial, for example, that silver chloride is insoluble and that a solution of a soluble silver salt added to a solution of a chloride will form a precipitate and by analysis we find this precipitate to be silver chloride. In working with chemicals one soon learns and easily remembers those things which will not remain in solution together. These elements are known in *Materia Medica* and prescription writing as *incompatibles*.

**Incompatibles.**—From what we have just learned it is obvious that silver nitrate and sodium chloride cannot be put in the same solution. Any salt of silver is said to be chemically *incompatible* with a chloride. The nurse, the pharmacist and the physician must remember the incompatible substances with which they have to work.

#### SUMMARY OF CHAPTER XIII.

By analysis with an electric current water is found to contain hydrogen and oxygen in the ratio of two

volumes H to one volume O. A mixture of  $2\text{H} + 1\text{O}$  can be exploded by a spark. All the gases disappear and water is formed (synthesis). Water therefore must consist of  $\text{H}_2\text{O}$  or some multiple of  $\text{H}_2\text{O}$ , as  $\text{H}_4\text{O}_2$ , etc. Molecular weight determinations of steam give approximately 18, which we find is the sum of the atomic weights of  $\text{H}_2$  and O, and not of  $\text{H}_4$  and  $\text{O}_2$ . Therefore the formula of water (vapor) is  $\text{H}_2\text{O}$ .

Water is a very stable compound.

Water combines with some oxides to form hydroxides ( $\text{Ca}(\text{OH})_2$ ,  $\text{NaOH}$ , etc.), and with  $\text{SO}_2$  or  $\text{SO}_3$  to form acids ( $\text{H}_2\text{SO}_3$  = sulphurous acid; or  $\text{H}_2\text{SO}_4$  = sulphuric acid).

An acid and a base combine to form a salt plus water.

The group OH is called hydroxyl.

Substances in solution are *ionized* or slightly dissociated. There is a constant and rapid interchange of ions until two ions, which, when united form an insoluble compound. These two ions and similar ions drop out of the field of interaction. Precipitation continues as long as both varieties of ion are present (*i. e.*, until one kind is used up).

Any two substances which cannot exist together in the same solution are said to be incompatible.

## CHAPTER XIV.

### HYDROGEN PEROXIDE.

WATER is the monoxide of hydrogen  $\text{H}_2\text{O}$ , while hydrogen peroxide is the dioxide  $\text{H}_2\text{O}_2$ . It is water plus one atom of oxygen.

**Preparation.**—Hydrogen peroxide does not occur in nature in any appreciable amounts as far as we know. It is perhaps in the atmosphere in very small quantities.

It is prepared by treating the peroxide of barium  $\text{BaO}_2$  with hydrochloric acid  $2\text{HCl}$ .

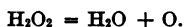


It is then purified by *fractional distillation* (see Chapter X). The ordinary solution on the market is a 3 per cent. solution in water, though a 30 per cent. solution can be purchased.

**Properties.**—Pure hydrogen peroxide is a clear colorless liquid very much like water but is one and one-half times heavier (specific gravity = 1.4996). It has no odor, possesses a peculiar, slightly acrid, taste and produces a soapy froth when taken into the mouth. It gives up the extra atom of O on standing, especially in a warm place and exposed to light. It should, therefore, be bottled in amber-colored glass and kept in the ice-box. The solutions of 3 per cent. hydrogen peroxide on the market generally contain a small

amount ( $\frac{1}{8}$  grain to the ounce) of acetanilid to preserve it. Old solutions should not be used. It is safer to replace the stopper with a cotton plug, which allows any liberated gas to escape.

**Uses.**—Hydrogen peroxide gives up its oxygen to oxidizable substances in the sense of the following equation:



It is therefore useful as an oxidizing agent. The 30 per cent. solution is rarely used in medicine; when dropped on the skin it is very caustic. The 3 per cent. solution is non-poisonous and is an antiseptic and for these reasons used in cleaning ulcers and infected wounds. On coming in contact with pus, blood or living tissues it gives up its oxygen which can act in *nascent* form to kill bacteria and destroy pus. By means of the bouyant effect of the gas it is useful in cleaning remote portions of a wound but should never be used in a deep or punctured wound because the pressure of the gas liberated may drive the infection further up in the tissues. It is useful in diphtheria and other throat infections and as a mouth wash.

Hydrogen peroxide is also capable of bringing about the reduction of compounds containing oxygen. The affinity of the atom of O in  $\text{H}_2\text{O}_2$  for other atoms of O is so strong that it unites with one atom to form a molecule of oxygen,  $\text{O}_2$ .

**Catalyzers.**—Substances like blood, and pus, etc., which decompose hydrogen peroxide are said to contain *catalases* (ferments), which bring about the liberation of the oxygen but do not enter into the

products of the reaction. A small amount of catalase will decompose an unlimited amount of  $\text{H}_2\text{O}_2$  and remain active. Platinum in a finely divided condition, will decompose hydrogen peroxide, and is an example of an inorganic catalyzer.

When zinc is acted upon by  $\text{HCl}$  with the formation of  $\text{H}_2$ , a small amount of stannous (tin), chloride or platinum chloride is put in to accelerate the reaction. It helps the reaction proceed but is not used up—a small amount is useful indefinitely.

Another example of a catalyzer which is made use of in ordinary life in platinized asbestos as a pocket lighter. Methyl (wood) alcohol in a metal vial, gives off fumes which come into contact with some platinized asbestos. Air is allowed to mix with it and acting under the influence of the finely divided platinum the oxygen begins to combine with the alcohol. This reaction generates heat until there is enough heat to ignite its alcohol fumes into a blaze.

#### SUMMARY OF CHAPTER XIV.

Hydrogen peroxide is water plus one atom of oxygen, that is,  $\text{H}_2\text{O}_2$ .

It is prepared by treating barium peroxide,  $\text{BaO}_2$  with  $\text{HCl}$  and distilling.

Hydrogen peroxide is an oxidizing agent, easily giving up the extra atom of oxygen in nascent form. It can also act as a reducing agent on compounds containing oxygen. It possesses germicidal power. The ordinary solution is 3 per cent. Stronger solutions are

dangerous. All solutions should be kept in a dark, cool place and preferably in a cotton-plugged bottle.

When hydrogen peroxide comes into contact with living tissues it is decomposed by certain ferments called *catalyzers*. There are also certain inorganic substances like platinum black, which act catalytically on  $H_2O_2$ .

In cleaning wounds with hydrogen peroxide solutions care should be exercised in its use. There is liability of forcing the infecting agents more deeply into the tissues. Hydrogen peroxide should not be employed in dressing deep and punctured wounds.

## CHAPTER XV.

### CHLORINE.

(At. wt. = 34.45.)

**Occurrence.**—Chlorine does not occur in nature free, but in very large amounts in combination with other elements. It is one of the chief elements in sea water, being found there combined with sodium, potassium, magnesium, and others. It occurs also throughout the crust of the earth in soluble and insoluble combinations. As soon as water comes in contact with the soluble salts they are dissolved and carried to the sea. About one-third of the salts of the Dead Sea is sodium chloride (table or common salt), that is, there are seven pounds of sodium chloride in every hundred pounds of the water.

Animal bodies contain chlorine, as chlorides in relatively large amounts. Blood contains about 0.8 parts sodium chloride per hundred. It is given off in the urine, sweat and feces in large amounts. Well waters contaminated with sewage even remotely will exhibit evidences of pollution by reason of high chlorine content.

**Preparation.**—The laboratory method of producing chlorine is to heat manganese dioxide with hydrochloric acid. The latter is oxidized to form water and sets free chlorine:





An easier method is to allow hydrochloric acid to act on bleaching powder.

Commercially chlorine is produced by passing an electric current through sea-water. The sodium chloride is broken up setting free chlorine and hydrogen, and leaving behind soda lye.

**Uses.**—Chlorine is a disinfectant and finds use as such in the cleaning of closets, drains, etc., and in the disinfection of excreta from patients having infectious diseases. It is a very active element when free, attacking metals and bleaching colors, and for these reasons must be used with care. It is applied in the form of *bleaching powder* (hypochlorite of lime, which see). In recent years chlorine has been used in the treatment of public water supplies to rid them of infectious organisms. One part of chlorine in 5,000,000 parts of water is sufficient. In this case also it is applied in the form of bleaching powder.

Chlorine water is found in the U. S. Pharmacopœia as *Liquor Chlorig Compositus*, but is very rarely used.

**Properties.**—Chlorine is a yellowish-green gas, soluble in water with the production of a yellow solution. It has a characteristic odor and is very irritating to mucous membranes when inhaled. It should be handled with great care and all experiments with it carried out under a hood with good ventilation. It is heavier than air and settles to the bottom of the hood.

Chlorine is very active chemically: it combines with almost all elements. Copper, iron, phosphorus, sodium, potassium, etc., burn in chlorine gas just as

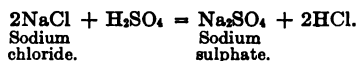
in oxygen. Most of those mentioned do not have to be ignited; for example, if copper foil is placed in chlorine gas it glows and  $\text{CuCl}_2$  is formed. Sodium and potassium unite with moist chlorine almost explosively forming the chlorides. As a general rule chlorides of metals with the exceptions of silver and lead are soluble in water.

**Hydrochloric Acid.**—Reference has already been made to the fact that hydrogen gas and chlorine gas combine volume for volume. If equal volumes of these gases in an absolutely dry state be placed together and an electric spark passed through, nothing will happen, but if a minute trace of moisture be present the mixture will explode, forming two volumes of hydrochloric acid gas. The formula for hydrochloric acid is  $\text{HCl}$ , or some multiple of this as  $\text{H}_2\text{Cl}_2$  or  $\text{H}_3\text{Cl}_3$ , etc., because they combine in equal volumes. If now we determine the molecular weight of hydrochloric acid we find it equal to 36.458. Since  $\text{H} = 1.008$  and  $\text{Cl} = 35.45$ , then  $\text{H}_1\text{Cl}_1$  must be the formula. If our formula were  $\text{H}_2\text{Cl}_2$  or more then the molecular weight determinations would give higher figures as 72.916 or 109.374, etc.

**Occurrence of  $\text{HCl}$ .**—Hydrochloric acid interests us because it occurs free in the stomach. It is secreted normally by glands in the stomach wall and is necessary in the peptic digestion of meats, eggs, etc.

**Preparation.**—Hydrochloric acid is produced on a commercial scale by allowing sulphuric acid to act on a chloride like sodium chloride (common salt). Sulphuric acid has a stronger affinity for a base than

hydrochloric and therefore displaces it in the sense of the following equation:



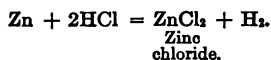
We shall later learn how sulphuric acid is produced.

**Uses.**—Hydrochloric acid is used extensively in the industries and in chemical analysis and synthesis. In medicine it is administered in dilute form to patients who have an insufficient secretion of acid in the stomach.

**Properties.**—As has been stated hydrochloric acid is a gas. We shall come to know it in water solution. Water will absorb the gas until a 39 per cent. solution is formed. This is the concentrated Hydrochloric Acid (sometimes called Muriatic Acid) of commerce. It has a pungent odor and its fumes strongly irritate mucous membranes. It burns the skin and destroys clothing.

The dilute hydrochloric acid of the Pharmacopœia contains 10 per cent. of the gas (1 part of concentrated hydrochloric acid plus 3 parts water).

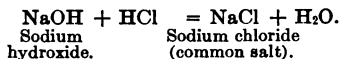
Hydrochloric acid attacks metals such as iron and zinc with the liberation of hydrogen and the formation of the chloride of the metal:



All the characteristics of an acid obtain in hydrochloric acid. It is in order then to discuss briefly the acids in general.

**Acids.**—To say that a substance is an acid is to say it is sour (Latin *acidus* = sour). Besides attacking

metals with the formation of hydrogen as indicated above, acids unite with bases (like soda lye) to form salts and water:



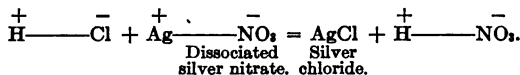
Acids turn blue litmus<sup>1</sup> red.

The sour taste of an acid has been found to be due to the ionized hydrogen. In water hydrochloric acid is dissociated into  $\overset{+}{\text{H}}-\overset{-}{\text{Cl}}$ . Vinegar is an acid (acetic acid) and its sourness is due likewise to the dissociated H (acetic acid in water =  $\text{CH}_3\overset{-}{\text{CO}_2}-\overset{+}{\text{H}}$ ).

Acids then are substances possessing a sour taste; chemically we say a compound possesses acid properties which, when dissolved in some dissociating solvent (like water), yields hydrogen ions.

**Test for Hydrochloric Acid.**—To detect hydrochloric acid in a solution, we place in it a piece of paper colored with blue litmus (litmus paper), and should it turn red we know that some sort of an acid is present. Having determined that some acid (an ionized H) is present we want to see whether chlorine is the other ion.

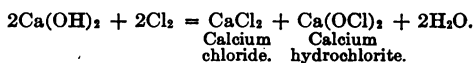
We remember from the paragraph on dissociation that when a Cl ion comes in contact with a silver ion a precipitate is formed. We therefore add some soluble salt of silver and in solution it is ionized.



<sup>1</sup> Litmus is a blue dyestuff obtained by fermenting certain coarsely ground lichens. It is used to test for acids, turning red when they are present and are turned back to blue by bases. When acids are completely and exactly neutralized by bases, the salts formed have no effect on red or blue litmus.

The AgCl comes down as a white precipitate which turns dark on standing. It is soluble in ammonia. We added silver nitrate and obtained a precipitate of silver chloride, therefore there must have been an ionized Cl present, and hydrochloric acid was in solution.

**Bleaching Powder.**—When chlorine is passed into slaked lime (calcium hydroxide  $\text{Ca}(\text{OH})_2$ ), calcium hypochlorite and calcium chloride are formed.



The mixture is put on the market as bleaching powder, so named on account of its ability to bleach. In slightly acid solutions chlorine is liberated. On account of the very great chemical activity of chlorine colors are bleached and bacteria killed. Reference has already been made (under chlorine) to the disinfecting power of this substance.

**Other Compounds of Chlorine.**—Chlorine combines with oxygen and with sulphur to form a large number of compounds but they are not of interest here. The chlorides of the metals, which are of interest here, will be discussed under the particular metals.

#### SUMMARY OF CHAPTER XV.

Chlorine occurs extensively in nature in common salt beds (sodium chloride) and sea water (chlorides of several metals). It is a very prominent and important mineral constituent of the animal body. Blood contains about 0.8 per cent. NaCl.

Chlorine is prepared by treating bleaching powder with hydrochloric acid or oxidizing HCl with  $\text{MnO}_2$ . Commercially it is produced by passing an electric current through brine.

Free chlorine is used as a bleaching and as a sterilizing agent. One part free chlorine in 5,000,000 parts of water kills disease-producing organisms like the bacillus of typhoid fever.

Chlorine is a yellowish-green gas with a characteristic odor. It is very active chemically. Metals like copper, iron, and sodium burn in it, forming chlorides.

Hydrogen unites with chlorine to form a very important compound, hydrochloric acid (HCl).

Hydrochloric acid is found normally in the human stomach where it is necessary for protein digestion. When there is not sufficient acid in the stomach HCl is administered by mouth in dilute form. The concentrated hydrochloric acid of commerce is a 39 per cent. solution of the gas in water. It is prepared by treating a chloride with concentrated sulphuric acid.

Hydrochloric acid attacks metals like zinc, forming the chloride of zinc ( $\text{ZnCl}_2$ ) and setting free hydrogen gas. (See Preparation of Hydrogen.)

The chemical definition of an acid is as follows: an acid is a compound of H with some radical or negative element which when dissolved in water yields hydrogen ions. Likewise, bases yield hydroxyl groups (OH) on dissociation. Acids and bases neutralize one another, producing salts.

The test for hydrochloric acid (or a chloride) is the addition of some soluble silver salt. If chlorides are

present a white precipitate of  $\text{AgCl}$  appears. This precipitate can be dissolved by adding a small amount of ammonia.

Bleaching powder is formed by passing chlorine gas over water-slaked lime. The mixture of calcium chloride and calcium hypochlorite formed constitute bleaching powder.

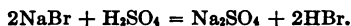
## CHAPTER XVI.

### BROMINE—IODINE—FLUORINE.

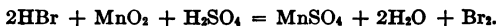
#### **BROMINE** (At. wt. = 80).

**Occurrence.**—Bromine, like chlorine, never occurs free in nature. It is found combined with metals forming the bromides in salt deposits and in the sea, though not in such quantities as the chlorides are found.

**Preparation.**—The bromine salts may be decomposed by an electric current to produce bromine, or sulphuric acid may be used to set free hydrobromic acid (just as hydrochloric acid is produced) thus:



and we may free the bromine of the hydrobromic acid from its H by allowing the H to combine with O through the agency of some oxidizing agents:



**Uses.**—Bromine itself finds very little use in medicine. It is used in the laboratory in one of the methods of urine analysis and is useful in synthetic chemistry. The salts, especially sodium, potassium, lithium, and strontium as nerve depressants.

**Properties.**—Bromine is a very heavy, dark, brownish-red, mobile liquid, evolving at ordinary temperatures reddish fumes, highly irritating to the eyes and mucous



membranes. Its odor is peculiar and penetrating, resembling chlorine.

Its chemical properties are very much like chlorine: in combination with the metals bromides are formed and in union with hydrogen hydrobromic acid results. The silver salt (AgBr) silver bromide is insoluble in water like silver chloride, so that the addition of a soluble silver salt is also a test for bromides. The appearance of the precipitate obtained (AgBr) is indistinguishable by inspection from the precipitate obtained with chlorides. Chemically we are able to detect the difference by adding a few drops of ammonia to the precipitate: if the precipitate is dissolved quickly we know it is silver chloride—if it is dissolved with difficulty it is silver bromide.

### IODINE (At. wt. = 127).

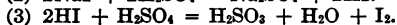
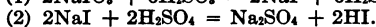
**Occurrence.**—Iodine is very widely distributed in nature but occurs in small quantities. It is found in small amounts with the chlorides and bromides in deposits and in the sea. It was first found in sea weed ash and has since been detected in other plants and in the lower and higher forms of animal life found in the sea. It is stated that a certain tropical sponge contains 14 per cent. (of the dry matter) iodine. In mammals it seems to play a very important role in the thyroid<sup>1</sup> gland since it is found there in considerable amount. (In the sheep over 9 per cent. of the dried

<sup>1</sup> The thyroid gland is found in the front of the neck. It is an enlargement of this gland that is called goitre.

gland is iodine.) White blood corpuscles also contain a very small amount of iodine in their composition.

Iodine is also found in minute quantities in the air and dust.

**Preparation.**—Most of the iodine is extracted as a sodium salt (sodium iodate,  $\text{NaIO}_3$ ) from the salt-petre deposits of Chile. Europe and Japan furnish perhaps a third of the world's production (750 tons per year). The iodate is freed of its oxygen by means of a reducing agent, such as sulphurous acid and the iodine set free by means of the sulphuric acid formed:



These reactions take place in the same operation, and it will be seen that 1 molecule of  $\text{H}_2\text{SO}_3$  is reformed, therefore in carrying out such a process only 5 parts  $\text{H}_2\text{SO}_3$  are needed instead of 6 parts as indicated in equation 1.

**Uses.**—Iodine was discovered a hundred years ago, but was not employed in medicine until 1831. Today it is used more widely in medicine and surgery than any other substance. It is a counter-irritant, parasiticide, absorbent, and alterative. It is widely used to disinfect wounds and is finding application more and more for preparing the site of operation and the operator's hands. It is more efficient when the tissues are dry.

Iodine is used in chemical synthesis and analysis, and in staining in the bacteriological laboratories.

In combination with hydrogen as hydriodic acid (HI) and as the iodides of sodium, potassium, and lithium, it is employed extensively in treatment. The world's consumption is about 750 tons per year.

**Properties.**—Iodine is a heavy blue-black solid, with a metallic luster. When heated these crystalline scales do not melt but vaporize immediately, giving off dark purple fumes. The vapor will deposit in needle crystals on coming into contact with a cool surface; in fact these properties are made use of in the purification of iodine. The process is called *sublimation*.

Iodine is soluble in alcohol and chloroform but very slightly soluble in water. When potassium iodide is present it will go into water solution easily, producing what is known in the laboratory as Lugol's solution.

Chemically, iodine is an oxidizing agent like chlorine and bromine; that is, in water solution it unites with the H of the water in the presence of a reducing agent and leaves the O of the water free to combine with this reducing agent. Thus sulphur dioxide ( $\text{SO}_2$ ), a reducing agent, is oxidized to sulphur trioxide ( $\text{SO}_3$ ) by iodine in the presence of water.

Silver iodide is insoluble in water and ammonia. It differs from the chloride and iodide in color; silver iodide is yellow while the other two silver salts are white.

Silver iodide is darkened on exposure to light though not as much as silver bromide. This change affected by light is the basis of photography, which will be discussed in the chapter on Silver and its Salts.

Iodine possesses the peculiar characteristic of combining with starch to produce a blue color which dis-

appears on boiling and returns on cooling. Iodine will produce the blue color only when it is uncombined; thus hydriodic acid or potassium iodide, etc., will not blue a solution of starch. This reaction may be used to test for either starch or iodine—if iodine is sought in a solution add starch solution—if starch be sought add a drop or two of Lugol's solution.

### FLUORINE (At. wt. = 19).

Fluorine is an active element belonging to the same group as chlorine, bromine, and iodine. It never occurs uncombined in nature. It is found as fluor spar (calcium fluoride) which is used as a flux in iron furnaces.

Fluorine is prepared in a manner similar to the method used for chlorine, and on account of its activity cannot be kept in glass but is bottled in paraffin. Its chief interest to us lies in the fact that the sodium salt is sometimes used unlawfully as a preservative in foods.

Hydrofluoric acid, HF, is used to etch glass.

**The Halogens.**—These four elements, chlorine, bromine, iodine, and fluorine form a group known as the *halogens* (salt-producing). They possess certain chemical and physical similarities of especial interest to the chemist, but have no application here.

### SUMMARY OF CHAPTER XVI.

Bromine is a very heavy, dark, brownish-red, mobile, caustic liquid. It never occurs free in nature. It is found as a salt (*e. g.*, NaBr) in the sea.

It is produced by passing an electric current through a solution of a bromide.

Bromine is little used in medicine but is employed extensively in synthetic chemistry. In its chemical properties bromine is very similar to chlorine.

Iodine is a heavy, blue-black solid with a metallic luster. It does not melt but vaporizes and can be sublimed. It is soluble in alcohol and chloroform. Very slightly soluble in water, though it will dissolve in a solution of KI in water (Lugol's solution).

Iodine is found as sodium and potassium iodate in Chile saltpetre deposits and in sea weeds. It is prepared by treating the iodate with sulphurous acid.

Iodine is perhaps the most widely used substance employed in medicine and surgery. It is employed as counter-irritant, parasiticide, absorbent, and alterative. Its chemical properties are similar to those possessed by chlorine and bromine though it is not so active chemically. Starch solutions turn blue in the presence of free iodine in cold solutions, on account of the formation of starch-iodide.

Fluorine is the most active of this group of halogens and must be kept in paraffin. On account of its power to attack glass HF is used in etching and forms the basis of diamond inks. The sodium salt is sometimes used illegally as a food preservative.

Chlorine, bromine, iodine and fluorine belong to a group known as the *halogens* (salt-producing).

## CHAPTER XVII.

### SULPHUR.

(At. wt. = 32.)

**Occurrence.**—Sulphur is found in a free state as brimstone in volcanic areas<sup>1</sup> and in combination in mineral deposits. Fool's gold which one sometimes sees in coal is a combination of sulphur and iron,  $\text{Fe}_2\text{S}_3$  (iron sulphide). Sulphides of copper, zinc, lead, etc., also occur in large quantities in the earth's crust. The so-called sulphur springs contain hydrogen sulphide,  $\text{H}_2\text{S}$ , in solution.

On exposure to oxygen and oxidizing agents hydrogen sulphide deposits sulphur:



Sulphur is purified by distillation. The chemically pure product is obtained by recrystallization from solution in carbon bisulphide.

**Uses.**—Sulphur itself is used in various skin diseases due to parasites and other causes. It is sometimes administered internally for its cathartic action, due, doubtless, to the sulphides formed in the intestinal tract, for sulphur itself is inert. Sulphur is used in gun powder, and in the manufacture of certain kinds of matches.

<sup>1</sup> Sulphur deposits are found in Louisiana in sufficient quantities to warrant mining.

Burning sulphur is used in fumigations where it is desired to kill insects or animals, such as rats. Sulphur dioxide gas results from the burning and this, while not bactericidal, is very efficient for ridding buildings and vessels of insects, etc. Sulphites and sulphuric acid are useful in many industries.

**Properties.**—Sulphur is a pale yellow solid, light in weight, melts at a little above boiling-point of water (at  $118^{\circ}$  C.) and boils at about  $450^{\circ}$  C. On being melted it becomes darker in color, gradually assuming a dark brown shade. The solid is crystalline or non-crystalline (amorphous). It is possible to obtain easily two different kinds of crystallized sulphur. Not only are the shapes of these two kinds of crystals different, but they are found to contain different amounts of energy.

**Chemical Properties.**—Sulphur is relatively inert at ordinary temperatures but on heating will combine with a large number of elements. It combines with hydrogen to form two compounds  $H_2S$  and  $H_2S_2$ . The former is far more important as it finds extensive use in the laboratory. By the use of hydrogen sulphide the various metals can be separated into groups on account of the differences in solubility of their sulphides. Every laboratory has a hydrogen sulphide generator. The odor of this gas is extremely disagreeable and in large quantities poisonous. At least a part of the odor of decomposing flesh is due to hydrogen sulphide.

Sulphur combines with oxygen to form a number of compounds, only two of which need be mentioned.

SO<sub>2</sub>, sulphur dioxide, combines with water to form sulphurous acid:



which acid combines with bases to form sulphites. The gas SO<sub>2</sub> is liquefied and put up in cylinders for disinfection purposes.

SO<sub>2</sub> can be oxidized to sulphur trioxide SO<sub>3</sub>, a liquid, which in the presence of water becomes H<sub>2</sub>SO<sub>4</sub>, sulphuric acid. H<sub>2</sub>SO<sub>4</sub> is used in the preparation of many compounds and as a dehydrating agent. It is used in storage batteries and in steel pickling processes.

**Valence.**—The student has doubtless wondered that two or more compounds of the same elements can exist and also that some compounds can hold more hydrogen in combination than others.

It was seen that chlorine held one hydrogen in chemical union (HCl) while oxygen held two (H<sub>2</sub>O). Bromine, iodine and fluorine held one only, but sulphur held two (H<sub>2</sub>S). Later it will be seen that phosphorus can hold three or five. Now we learn that sulphur can hold two atoms of oxygen or three atoms, and since 1O is equivalent to 2H we have sulphur capable of holding an equivalent of 2H, of 4H or of 6H. This property of holding an element in combination is called *valence*. We accept hydrogen as the standard 1—then since chlorine can hold one H, the valence of chlorine is also 1; likewise the valence of O = 2.<sup>1</sup>

<sup>1</sup> These valences have been found experimentally and we simply remember them after we become accustomed to dealing with chemical substances. The student who wishes to know more of chemistry will find a discussion of the Periodic Law of Mendeléeff (in any thorough treatise on Chemistry) very interesting. We are able to predict what the valence will be in some cases before an element is discovered.



Some elements instead of having a non-variable valence may be able to hold different hydrogen equivalents according to the conditions. Chlorine usually has a valence of 1, but sometimes it may be 3 or 5, while sulphur is 2, 4, 6 or even 8.

The question of valence is very important in the study of the carbon compounds, so-called *organic* chemistry to be discussed presently.

#### SUMMARY OF CHAPTER XVII.

Sulphur is a pale yellow solid, melting at  $118^{\circ}$  C. and vaporizing at  $450^{\circ}$  C. It may be amorphous (non-crystalline) or crystallize in two forms.

It is found free in nature in volcanic deposits, and purified by distillation or crystallization out of carbon bisulphide.

Sulphur is not very active chemically. It can be burned to the dioxide,  $\text{SO}_2$ , or trioxide  $\text{SO}_3$ . These compounds are soluble in water, forming respectively, sulphurous ( $\text{H}_2\text{SO}_3$ ) and sulphuric ( $\text{H}_2\text{SO}_4$ ) acids.

Sulphur unites with hydrogen and metals to form sulphides like  $\text{H}_2\text{S}$ ;  $\text{Na}_2\text{S}$ , etc. It is applied in skin diseases and burned to rid ships, houses, etc., of insects and rats. Burning sulphur is not advisable for fumigation after contagious diseases, as  $\text{SO}_2$  has little germicidal action.

The power to hold atoms in chemical combination is spoken of as *valence*. The valence of H = 1; of Cl, Na, K, Br, I, etc. = 1, while the valence of O, S, Mg, Pb, etc. = 2. An element may possess different

valences as S, the valence of sulphur may be 2, 4, 6, etc., of Hg may be 1 or 2 ( $\text{HgCl}$  or  $\text{HgCl}_2$ ). There is no rule about valence except the place of the element in the periodic system and this is complicated. We must simply remember the valences of the elements commonly dealt with.

U of M

## CHAPTER XVIII.

### SODIUM, Na.

(At. wt. = 23.)

IN daily life and in the hospital we come into contact more often with compounds of this element than any other. Sodium is the metal of which common salt is the chloride and it is the basic element in soda lye and washing and in cooking soda.

**Occurrence.**—Sodium, on account of its chemical activity is never found free in nature. In combination with other elements it is found everywhere. The sea contains large amounts of the chloride; great mines of the chloride are found in Germany and as nitrate it forms the well known Chili saltpetre beds. Dust contains it in detectable amounts and even the dust-free atmosphere of the midocean holds it in small quantities.

Sodium chloride occurs in all the tissues of the body and in the blood; there are about eight parts in every thousand.

In biblical times sodium chloride was mined from deposits containing other salts. The expression "Salt has lost its savor" refers to lumps from which all the sodium chloride was taken away and some other salt left. It is impossible for pure sodium chloride to lose its savor unless changed chemically.

188

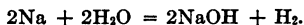
**Preparation**—The chloride and nitrate found in nature may be purified by crystallization, and the element itself is obtained by passing strong electric currents through molten soda lye (NaOH).

**Uses**.—The *saline* solution of the hospital is a solution of sodium chloride. It is an essential component of media for growing microorganisms. There are over twenty-five salts of sodium used in medicine.

As lye, extensive use is made of sodium in various industries. It is being used more and more in dye and other chemical syntheses.

Soap is sodium combined with acids from fats as we shall see when we study fats. Soda is necessary in glass manufacture—the silicate of sodium is one of the chief constituents of this useful substance.

**Properties**.—For a long time soda lye was thought to be an element, but when Sir Humphry Davy passed an electric current through it a bright metallic substance rose to the top and took fire when it came in contact with air. This was metallic sodium, a soft, steel-gray solid, which rapidly turns dark in air if a trace of moisture is present. Sodium is so active that it reacts with explosive violence if water is poured on it. Sodium hydroxide is formed and hydrogen is given off:



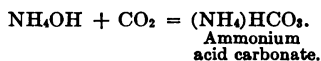
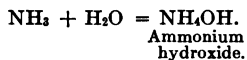
The heat generated by the reaction may be sufficient to ignite the hydrogen.

Sodium combines with chlorine in the presence of moisture to form sodium chloride (common salt).

It is obtained from sea-water by evaporation or mined from the salt deposits and purified by recrystallization. It crystallizes in characteristic hopper-shaped cubes which are hollow in the centre. These crystals melt at  $780^{\circ}$ , and fly to pieces (decrepitate) when heated. It is soluble in water to the extent of 36 parts per 100; *i. e.*, 100 parts saturated solution of sodium chloride contains 36 parts of the salt. Heat increases the solubility very little.

Sodium chloride is necessary for bodily function, and therefore a necessary constituent of the food. In flesh food it is present in sufficient quantities to balance the potassium, calcium, and magnesium salts, therefore carnivorous animals do not require salt. On the other hand, plants contain more potassium than sodium, and herbivorous animals seek *salt licks* in order to maintain the balance. Man subsisting on a mixed diet requires salt. Too much salty food without organic acids (vinegar, fruit juices) has the tendency to cause scurvy.

Sodium chloride is the starting-point in the manufacture of sodium salts. If ammonia ( $\text{NH}_3$ ) is added to a saturated solution of sodium chloride and carbon dioxide ( $\text{CO}_2$ ) passed through it at first ammonium carbonate is formed:



Now in the presence of NaCl there is a travel of ions, so we have a formation of ammonium chloride ( $\text{NH}_4\text{Cl}$ ) and of sodium acid carbonate, and since sodium acid carbonate is insoluble in the already saturated solution of salts it is precipitated:



The sodium acid carbonate is filtered off and purified by recrystallization.

**Cooking Soda.**—The sodium acid carbonate just described above is what we know as cooking soda. When an acid is added to it carbon dioxide gas is given off:



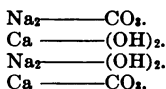
This is the principle of the raising of dough—the gas given off expands and makes the bread light. In practice hydrochloric acid is not used because: (1) the reaction would take place before the bread is heated and all the gas would come out of the dough leaving it flat; (2) the strong acid would affect the other ingredients; and (3) the resulting compound, NaCl, would make the bread too salty. For these reasons a weak acid is used so that the reaction does not take place until the dough becomes thick enough to hold the gas. An organic acid like lactic (the acid found in sour milk) or tartaric (manufactured from grapes) is used. The latter is mixed in powdered form with the sodium acid carbonate and sold as *baking powder*. On heating in the presence of moisture the tartaric acid combines with the sodium acid carbonate to form

sodium tartrate (Rochelle salt) and  $\text{CO}_2$  is given off.<sup>1</sup> Sodium acid carbonate is called sodium bicarbonate, for when it is heated to a high degree it gives off one part of  $\text{CO}_2$  and sodium carbonate  $\text{Na}_2\text{CO}_3$  remains:



This carbonate is known as *washing soda*. It has more basic properties than the bicarbonate. This salt crystallizes easily and depending upon conditions combines with one, seven or ten parts of water; that is, we obtain the compounds  $\text{Na}_2\text{CO}_3 \cdot \text{H}_2\text{O}$ ;  $\text{Na}_2\text{CO}_3 \cdot 7\text{H}_2\text{O}$  and  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ . These molecules of water are not simply occluded within the crystal but are chemically a part of it though it can be driven off by heat just as the  $\text{CO}_2$  is driven off from the molecule of the bicarbonate.<sup>2</sup> Sodium carbonate can be obtained also by leaching the ashes of sea weeds.

**Sodium Hydroxide.**—If sodium carbonate is treated with lime water  $\text{Ca}(\text{OH})_2$ , the interchange of ions would produce the following compounds:

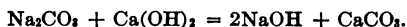


Now,  $\text{Ca---CO}_3$  is an insoluble compound, so that it is immediately precipitated and becomes  $\text{CaCO}_3$ .  $\text{Na}_2(\text{OH})_2$  is soluble and resolves into the normal

<sup>1</sup> This same reaction takes place when Seidlitz powder solutions are poured together, but here the substances being in solution the reaction takes place more vigorously.

<sup>2</sup> These molecules of water are known as *water of crystallization*. This same phenomenon is observed in many other compounds.

compound Na—OH. The whole reaction is expressed by the following equation:



The mother liquor after filtration is evaporated and the soda lye (caustic soda or sodium hydroxide, NaOH) is left in white, crystalline lumps. From sodium hydroxide the various salts of sodium are easily made.

#### SUMMARY OF CHAPTER XVIII.

Sodium is a soft, steel-gray solid which rapidly tarnishes in air containing even a trace of moisture. A small amount placed on water reacts violently to form sodium hydroxide (caustic soda) and sets free hydrogen. The hydrogen soon catches fire on account of the heat of the reaction.

Hydrochloric acid added to sodium hydroxide forms sodium chloride (common table salt), plus water. Sodium is found everywhere. Even dust contains detectable amounts of this element. Sodium chloride and nitrate occur in large salt beds and the former is the principle salt of sea-water. Sodium chloride is essential to the animal organism.

Cooking soda is sodium bicarbonate. Washing soda is more strongly alkaline and is the normal carbonate. Baking powder is a mixture of sodium bicarbonate plus some organic acid. When heated these two substances combine to form a salt and liberate  $\text{CO}_2$  which "raises" the dough.

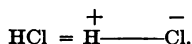
Sodium salts are characterized by their easy solubility in water. A very minute portion of sodium or a sodium salt imparts an intense yellow color to a flame.



## CHAPTER XIX.

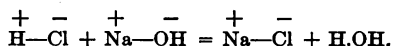
### ACIDS AND BASES—POTASSIUM.

WE have learned that an acid is a sour tasting substance which dissociates in solution into H ions and some other ions. An example of this is hydrochloric acid which gives rise to free H ions and free Cl ions thus:



We come to learn now that a base is a substance which in solution is dissociated, giving rise to free hydroxyl (OH) ions, as sodium hydroxide Na—OH.

When both an acid and a base are in the same solution they neutralize one another



As long as the acid and base are present in such amounts that there are as many H ions as there are OH ions they exactly neutralize one another and the solution is neutral.

**Molecular Solutions.**—How do we prepare solutions containing equal numbers of H and OH ions?

Let us take as a standard 1 gram of H in a liter of water. But the H is bound up with chlorine if we use hydrochloric acid. Therefore we use such an amount of hydrochloric acid as contains exactly 1 gram H.

The atomic weight of  $H=1$ , and of chlorine 35.45, making the molecular weight of  $HCl = 36.458$ . Therefore, in 36.45 grams  $HCl$  there is contained 1 gram  $H$ . Then 36.45 grams  $HCl$  in 1 liter is our standard solution which we shall term a *normal* solution.

To prepare a solution of a base to exactly neutralize our normal acid solution we must have enough  $OH$  ions in solution to be equivalent to 1 gram  $H$  in a liter. At. wt. of  $O = 16$  and of  $H=1$ , then mol. wt. of  $OH$  ion  $=17$ , and we must have, therefore, 17 grams  $OH$  to be equivalent to 1 gram  $H$ . The at. wt. of  $Na=23$ ,  $OH = 17$ , mol. wt.  $NaOH = 40$ . Therefore, in every 40 grams  $NaOH$  we have 17 grams  $OH$  or sufficient amount to neutralize 1 gram  $H$  ions. To make a normal solution of the base  $NaOH$  then we dissolve 40 grams  $NaOH$  in 1 liter of water. One liter containing 36.45 grams  $HCl$  will exactly neutralize 1 liter containing 40 grams  $NaOH$ .

From the above we learn then that a normal solution of an acid is one which contains 1 gram  $H$  ions per liter and a normal solution of a base contains 1 gram  $OH$  ions per liter. In the two instances cited above the normal solution corresponded exactly with a molecular solution (the molecular weight in grams dissolved in a liter).

Suppose we attempt to make a normal solution of sulphuric acid. The formula is  $H_2SO_4$  and it dissociates into



that is, two H ions for every molecule. Then a normal solution would be only one-half the molecular solution. An example of a base in which the same is true is calcium hydroxide<sup>1</sup> (slaked lime)  $\text{Ca}(\text{OH})_2$ , which dissociates into



A normal solution is indicated as follows: N/1, twice normal 2N and half-normal N/2. Molecular solutions are labelled M/1, etc., and called molar, half-molar, etc.

A chemically normal solution must not be confused with the so-called *physiological* normal saline solution. The latter is a misnomer and instead of being called *normal salt* solution it should be called M/8 (eighth molar), for this strength of solution (0.9 per cent. NaCl) has approximately the same osmotic pressure as the blood.

**Indicators.**—When acids (which see) were described it was stated that litmus turns red in the presence of acids and blue in the presence of bases. Therefore we have in litmus an indicator which when added to a solution tells us whether the solution is of acid reaction or basic (alkaline) reaction. Paper impregnated with litmus is used in urine analysis to ascertain whether the urine is acid or basic. The quickness with which the strip of paper turns indicates the relative strength of the acid or base though only to an approximate

<sup>1</sup> Let the student give directions for preparation of N/1  $\text{HNO}_3$  and N/1 KOH.

degree. We say then a sample of urine is strongly or weakly acid or alkaline (basic).

**Volumetric Analysis.**—Suppose we have a sample of vinegar in which we want to determine the amount of acid. If the vinegar is highly colored we may dilute it and add our indicator. Now we add slowly from a graduated tube (burette) a normal solution of a base until the indicator tells us that the vinegar is now no longer acid but neutral and 1 drop of base added brings about the alkaline color. The basic solution being standard (normal) we know how much base is in a liter and any part of a liter. Also, we can calculate how much acid will be neutralized by any part of a liter of the normal solution of base. Therefore, reading from the burette the number of cubic centimeters of base used to neutralize the acid we can calculate how much acid is in the vinegar. This method of determining the strength of a solution by testing it against a normal solution is called *titration*. Titration is the foundation of *volumetric analysis*. (For further details consult Sutton's Volumetric Analysis.)

There are other indicators besides litmus, some of which are acid and some basic, useful according to their particular properties. The nurse will employ, besides litmus, probably only phenolphthalein, alizarin, and Töpfer's reagent (dimethyl-amino-azo-benzene).

### POTASSIUM, K (At. wt. = 39).

Potassium is very similar to sodium. It reacts in the same manner and what was said of sodium in general

is true of potassium. It is found as potash (KOH) in ordinary wood ashes and the nitrate is found in large deposits (saltpetre beds). It is found distributed in the human body just as sodium is, though in smaller amounts. Potassium salts are used in medicine like sodium salts; practically the only difference is that they have a more depressing effect on the heart than the latter.

Lithium is another metal included in this group with sodium and potassium, called the alkali metals. Little use is made of its salts in modern medicine.

#### SUMMARY OF CHAPTER XIX.

An acid in solution is dissociated into H ion + a negative ion (*e. g.*, Cl). A base in solution is dissociated into a metal (let M represent any metal) + OH ion (hydroxyl).

The molecular weight of HCl is 36.458 (H=1.008; Cl=35.45). In 36.458 grams HCl there is 1.008 grams H (approximately 1 gram). A *normal* solution of an acid is a solution of such an amount of the acid as would represent 1 gram of ionized H per liter, in the case of HCl 36.458 grams per liter. Since masses of acids and bases react molecule for molecule instead of gram for grams, the amount of base necessary to neutralize the molecular weight in grams of HCl, would be the molecular weight of the base in grams (NaOH=40).

A normal solution of a base is such an amount of base as would represent 17 grams (sufficient to combine with 1 gram H) of hydroxyl in a liter of water.

Some acids like  $\text{H}_2\text{SO}_4$  possess *two* available H ions for every molecule—therefore a normal solution would be *half* the gram-molecular weight per liter. This is true also of dihydroxy bases like  $\text{Ca}(\text{OH})_2$ .

The term *normal* in speaking of solutions should be confined to strictly chemical meaning. Physiological salt solution is eighth molecular, M/8.

Certain dyes like litmus are colored differently by acids and alkalies (bases). The change from a basic to acid reaction is accompanied by a sharp change in the color of the solutions of the dye. This color change is made use of in testing acid solutions against bases for their relative strengths. Such testing or titration (titre = to test) is known as volumetric analysis. The properties and reactions of potassium resemble those of sodium.

## CHAPTER XX.

### PHOSPHORUS—ARSENIC—ANTIMONY— BISMUTH.

#### PHOSPHORUS (At. wt.=31).

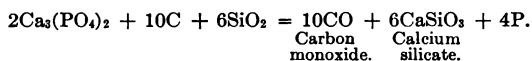
THE story of phosphorus forms one of the most interesting chapters in chemistry. Its occurrence in the body, its change of form and varied chemical activity bring it into prominence.

**Occurrence.**—Phosphorus occurs as the phosphate of calcium  $\text{Ca}_3(\text{PO}_4)_2$  in large deposits in the earth and is scattered through the soil also in the form of salts. It does not occur in the free state.

The mineral matter of bones is largely calcium phosphate even to such an extent that bone ash is used as a source for the preparation of phosphorus. The tissues of plants and animals also contain phosphorus in various combinations.

**Preparation.**—Calcium phosphate, sand (silicon dioxide) and charcoal (carbon) are heated in an electric furnace to a high temperature and the free phosphorus is given off. It is condensed and purified by redistillation.

The reaction which takes place is an interchange of silicate and phosphate and a reduction of the phosphoric acid by the carbon:



**Uses.**—Until recently phosphorus was employed in the manufacture of matches. On account of its poisonous effects on workmen its use now is forbidden by law. Phosphorus is used in rat poison and vermin-killer. In rare instances it is sometimes administered as a therapeutic agent. The hypophosphites were formerly administered as tonics though now it is believed that such treatment is worthless. The pentoxide ( $P_2O_5$ ) is a strong dehydrating agent and the chlorides  $PCl_3$  and  $PCl_5$  are used in building up certain organic (carbon-hydrogen) compounds.

**Properties.**—Cold phosphorus is a yellowish, brittle solid. On being warmed to room temperature it becomes soft and waxy and melts at a little above body temperature ( $45^\circ C.$ ). The free elements may exist in four different states: as *yellow phosphorus* (ordinary form); *red phosphorus*; *black phosphorus*, and *white phosphorus*. The first two are more important from a chemical stand-point and more is known of them.

*Yellow phosphorus* is so active that it must be kept under water. When oxygen comes in contact with it, some of the oxygen is converted into ozone, and there is also a chemical union of phosphorus and oxygen in which the oxide is formed, heat is liberated and the phosphorus glows. Undoubtedly the property which phosphorus possesses of glowing in the dark is in some way associated with oxidation. When the temperature reaches  $50^\circ C.$ , the phosphorus in contact with oxygen takes fire. Yellow phosphorus is very active chemically, forming some of the compounds which will later



be discussed. If yellow phosphorus is heated in some inert gas like nitrogen to  $250^{\circ}$  it is changed to a red amorphous powder known as *red phosphorus*. In this form phosphorus is not so active. It does not combine readily with other elements and does not take fire when heated in the presence of oxygen. It is much less poisonous than yellow phosphorus.

Yellow phosphorus combines with hydrogen to form phosphine  $\text{PH}_3$ , a gas, and  $\text{P}_2\text{H}_4$ , a solid. It also combines with oxygen to form oxides  $\text{P}_2\text{O}_3$  and  $\text{P}_2\text{O}_5$ , and with hydrogen and oxygen to form acids. Phosphoric acid  $\text{H}_3\text{PO}_4$ , which combines with bases to form phosphates and hypophosphorous acid  $\text{H}_3\text{PO}_2$ , which in like manner forms hypophosphites, are the more important.

A group of important substances found in yolk of eggs, in milk and in brain tissues, known as lecithins contain phosphoric acid, combined with acids from fats and an organic base.

### ARSENIC (At. wt. = 75).

Arsenic in some cases acts like phosphorus and in others like sulphur. It occurs in combinations in nature, never free. Its chief source is the iron compound  $\text{Fe}_2\text{As}_3$ , and it also occurs in a compound similar to Fool's gold (pyrites).

Arsenic is a gray, hard, brittle metal. It combines with various elements: hydrogen, oxygen, sulphur, the halogens and the metals. Its hydrogen compound reminds us of phosphine. Arsine  $\text{AsH}_3$  is formed when

hydrogen is generated in a solution containing arsenic. This is the basis of the well-known Marsh test for arsenic. The solution under suspicion as containing arsenic is allowed to run into a flask where hydrogen is being produced by the action of sulphuric acid on zinc. The nascent hydrogen reacts with the arsenic to form arsine which comes out the delivery tubes with the excess of hydrogen. The jet from the delivery tube is lighted and a cold porcelain dish held in the small flame. If arsine is present a metallic film of arsenic is deposited on the dish.

Fowler's solution is a solution of potassium arsenite:  $K_2HAsO_3$  formed by boiling arsenic trioxide with potassium acid carbonate.



Arsenic has come into prominence lately on account of its part in the composition of salvarsan which is an elaborated arsenic and benzene compound.

Arsenic is a common impurity of mineral acids and certain salts. Special methods are necessary to rid them of the last traces of this metal and the standards set by the Pharmacopœia allow only minute amounts of arsenic as impurity.

#### **ANTIMONY, Sb. (At. wt. = 120).**

Antimony is very much like arsenic: one seldom attempts to remember its properties except that they are almost the same as arsenic. Even the Marsh test with slight variation is used as a test for antimony.

Our interest in this metal lies in the fact that it is one of the chief components of tartar emetic which is potassium and antimony tartrate. Tartar emetic is put in compound syrup of squills.

Antimony is useful in making the alloy used for manufacturing type.

**BISMUTH, Bi. (At. wt. = 208).**

Bismuth is also like arsenic and closely allied in properties to phosphorus and antimony. It is not so active chemically as the other members of the group. In fact it is found free in nature. Bismuth is a crystallized solid but has not as much of the metallic sheen as arsenic and antimony.

Various salts exist. Their formulas are easily predicted when we know that bismuth is trivalent, that is, one atom will hold in combination three atoms of a monovalent element like chlorine. Bismuth hydroxide is  $\text{Bi}(\text{OH})_3$  and the nitrate is  $\text{Bi}(\text{NO}_3)_3$ . There is a combination of these two salts called the subnitrate  $\text{Bi}(\text{OH})_2\text{NO}_3$  in which it will be seen that the bismuth is not entirely nitrated but that two of the nitrate groups ( $\text{NO}_3$ ) are replaced by hydroxyl groups ( $\text{OH}$ ). On account of these two hydroxyl groups in the molecule the salt will react basic (alkaline) and is called the basic nitrate.

Other metals are capable of forming subnitrates (and other basic salts) also. Bismuth salts are used as astringents and in *x*-ray work. If the salts are taken into the intestine and an *x*-ray made it is found that

the rays are obstructed by the bismuth and shadows are cast on the plate. In this manner the movements and the shape of the stomach and intestines are studied. Sinuses may be studied in like manner.

### SUMMARY OF CHAPTER XX.

Phosphorus is a yellowish, brittle solid which becomes waxy on heating. It may exist in four forms: yellow, red, black or white. Yellow phosphorus is the most common. This variety is the most active chemically. It glows in air and may burn spontaneously if not covered with water. Yellow phosphorus is changed into an inert form (red phosphorus) on heating in an inert gas like nitrogen.

Phosphorus occurs as calcium phosphate in mineral deposits and in this form also constitutes the chief inorganic part of bone. The element is obtained from this salt by heating it in an electric furnace with a mixture of sand and charcoal.

Phosphorus combines with hydrogen to form phosphine  $\text{PH}_3$  and with oxygen and water to form several phosphoric acids. The salts of these acids are called phosphites, hypophosphites, phosphates, etc.

An important class of foodstuffs, lipoids (fat-like substances) may contain phosphorus in combination. Lecithin found in eggs, milk, and brain tissue is a phosphorus containing lipid.

Arsenic is very similar to phosphorus in its compounds. It is less active chemically and the free element possesses more metallic properties. It forms arsine with hydrogen (similar to phosphine) and also

forms acids and salts, arsenites, arsenates, analogous to phosphorus compounds. The Marsh test depends upon the formation of arsine,  $\text{AsH}_3$ . Arsenic compounds produce acute and chronic stages of poisoning. Some compounds, potassium arsenite, are administered, and many of the newer products of chemotherapy contained arsenic combined with organic radicals (salvarsan).

Antimony is similar to phosphorus and arsenic. Potassium and antimony tartrate is the principal medicinal preparation. Bismuth belongs to the same chemical group. It is a metal, sometimes found free in nature. The element is trivalent, *i. e.*, holds in combination three chlorine atoms or hydroxyl groups ( $\text{BiCl}_3$ ,  $\text{Bi(OH)}_3$ ). The nitrate  $\text{Bi(NO}_3)_3$  and especially the subnitrate  $\text{Bi(OH)(NO}_3)_2$ , also the sub-gallate are used in local applications for their astringent effect. Bismuth salts are relatively impervious to Röntgen rays and, hence, find use in röntgenography.

## CHAPTER XXI.

### CALCIUM.

(At. wt. = 40.)

SODIUM, potassium and lithium are called the alkali metals on account of their ability to form bases (alkalies). Calcium, strontium and barium also possess basic qualities but not to the same extent. This group is called the alkaline earths. Calcium will be discussed as the most interesting representative of the group. Barium is used little in medicine, and only one salt of strontium need be known, viz., strontium bromide.

**Occurrence.**—Limestone and chalk are impure calcium carbonate. Marble is a pure crystallized limestone. The calcium phosphate beds have been mentioned under phosphorus and fluorite ( $\text{CaF}_2$ ) under fluorine.

**Lime.**—It has already been stated (Chapter I) that lime results from the heating of limestone. The reaction is very simple, carbon dioxide is given off.



If the gas be kept confined in the chamber with the lime it will recombine on cooling. This is one of the simplest examples of a reversible reaction.

When water is poured on lime slaking takes place with the evolution of heat.



In solution calcium hydroxide dissociates into  $\text{Ca}^{++}$ ,  $\text{OH}^-$ ,  $\text{OH}^-$  and is therefore basic.

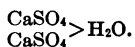
**Bleaching Powder.**—Reference to bleaching powder has already been made in connection with chlorine. Chlorine gas passed over moist lime forms equal parts of  $\text{CaCl}_2$  and calcium hypochlorite  $\text{Ca(OCl)}_2$  or one compound:



According to Jones it is the latter toward which most things point.

Whatever the exact composition of bleaching powder may be, it is the most convenient means of transporting chlorine, for, all the chlorine which goes into the reaction is recovered when the bleaching solution is rendered acid. Even a weak acid like carbonic suffices to replace the chlorine forming calcium carbonate. For this reason bleaching powder exposed to the air over *long* periods loses its characteristics because of the  $\text{CO}_2$  in the air.

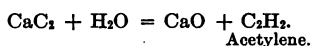
**Plaster of Paris.**—Calcium sulphate crystallizes with two parts of water of crystallization,  $\text{CaSO}_4 \cdot 2 \text{H}_2\text{O}$ . If the crystals are heated slightly above the boiling point of water ( $107^\circ$ ) only a part of the water is lost forming the compound



This is plaster of Paris. When water is added the compound  $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$  is rebuilt and the crystals fuse together and harden (set) into a mass. If all the water is driven off in the preparation of plaster of Paris (*i. e.*, heated too high) the hydration (adding of water) takes place too slowly to be of use. It will be remembered that plaster of Paris is used in the making of splints. The student now understands why it must be worked quickly after the water is added.

**Calcium Chloride.**—Calcium chloride has a great attraction for water and is therefore useful to the chemist as a dehydrating agent. Salts absorbing water on exposure to the air are said to be *hygroscopic*.

**Acetylene.**—Acetylene lamps are well known. Acetylene is a compound of hydrogen and carbon which burns with a brilliant light. Acetylene is manufactured by allowing water to come into contact with calcium carbide  $\text{CaC}_2$ .



Calcium carbide is made by heating a mixture of lime and powdered charcoal in an electric furnace.

**Flame Tests.**—A small amount of a calcium salt introduced into the flame of a Bunsen burner produces a dark red color. Strontium salts impart a brilliant red and barium salts a green color to the flame.

## SUMMARY OF CHAPTER XXI.

Calcium, strontium and barium have certain characteristics in common: their carbonates and sulphates



are relatively less soluble in water and all three form hydroxides which are more or less alkaline. These elements are bivalent and constitute a group in the periodic system known as the alkaline earths. The hydroxides of the alkaline earths are not so strongly basic in character as the alkali metal hydroxides.

Calcium is the most important member of the group. It does not occur free in nature, but its salts especially the carbonate and phosphate are very common. Marble is crystallized calcium carbonate. Lime is  $\text{CaO}$ . It results from the intense heating of  $\text{CaCO}_3$ . Water slaking means the addition of water to  $\text{CaO}$ , forming  $\text{Ca(OH)}_2$  (calcium hydrate).

Bleaching powder



is formed by passing chlorine gas over lime. It is the most economical and convenient method of transporting chlorine for bleaching and disinfecting purposes. Acids liberate chlorine from bleaching powder.

*Plaster of Paris* is calcium sulphate only partially hydrated. When water is added to a paste of plaster of Paris full hydration quickly results and calcium sulphate crystals ( $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ ) form (setting).

Salts like calcium chloride which absorb water from the air are said to be hygroscopic.

Acetylene is formed by the action of water on calcium carbide.

Calcium salts impart a dull red, strontium salts a brilliant red, and barium salts a green color to the Bunsen flame.

## CHAPTER XXII.

### MAGNESIUM GROUP.

#### **MAGNESIUM (At. wt. = 25).**

MAGNESIUM, zinc, mercury, cadmium and glucinum belong to this group. Magnesium is of greatest importance, especially from the stand-point of therapeutics. It occurs in nature as the carbonate, the chloride, the silicate, and the sulphate. One of the chief constituents of asbestos is magnesium silicate. Epsom salt is magnesium sulphate. The element magnesium, which can be obtained by electrolysis of its salts is a white, very light metal. It burns in air with a brilliant, white light, and for this reason is used extensively in fire-works.

**The Flame Test.**—Just as metallic magnesium burns characteristically in air other metals like sodium and potassium can be identified by their action in a flame. Sodium burns with a bright, yellow light and potassium gives the flame a peach-blossom color. Many metals, for example, iron, will not burn unless powdered and dusted into a flame, but if a piece be heated white hot and placed in an atmosphere of pure oxygen it will burn brilliantly. In testing the various elements it is not essential that they be in the free state. For example, any salt of sodium will color a flame yellow;

potassium salts also yield the characteristic peach-blossom color, strontium salts impart brilliant red; calcium a dull red, barium and copper, green, to the flame. If the light from the flame in which any salt is placed is conducted through a prism in such a manner as to resolve it into its primary colors characteristic bands of different colors will appear. This is known as spectrum<sup>1</sup> analysis. By these means the various elements have been detected in the sun and stars.

**Magnesia.**—When metallic magnesium is burned magnesium oxide,  $\text{MgO}$ , is formed. This compound is what is known as magnesia. It is insoluble in water, but when mixed with water to form a cream it forms the well-known milk of magnesia.

Magnesia is generally made by heating the carbonate. The  $\text{CO}_2$  is driven off leaving  $\text{MgO}$ .



Problem for student: What is the chemical reaction when milk of magnesia is taken into the stomach where free hydrochloric exists?  $\text{MgO} + 2\text{HCl} = ?$

Why do you write two parts  $\text{HCl}$ ?

**Epsom Salt.**—Epsom salt as found in the Epsom springs, is magnesium sulphate. It crystallizes with seven parts water of crystallization,  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ . There is a mineral consisting of  $\text{MgSO}_4 \cdot \text{H}_2\text{O}$ .

<sup>1</sup> A beam of white light passing through a prism is spread out into the colors of the rainbow—violet, indigo blue, green, yellow, orange, and red. All these colors together are called the spectrum, and the instrument for observing this phenomenon and the various bands formed by different elements is called a spectroscope. The elements in a molten condition produce characteristic lines in the spectrum.

When magnesium sulphate is dissolved in water it absorbs heat and makes the solution cold, thus reducing the solubility. In making strong solutions, therefore, warm water should be used.

**MERCURY, Hg. (At. wt. = 200).**

Mercury is the only metal which exists in the liquid state at ordinary temperatures. It occurs free in nature and is separated from the substances with which it is found by distillation.

On account of the fact that it solidifies at a very low temperature, and also that its volume is changed to a considerable extent when its temperature is varied, it is very useful for making thermometers.

**Amalgams.**—When mercury and gold come into contact they combine—the gold loses its yellow color and seems to be silvered over. If there is an excess of mercury the gold will dissolve in it. The same is true of silver, magnesium, calcium, and other metals. The combination of any of these metals with mercury is called an amalgam. Sodium and potassium form amalgams with mercury which are solid at ordinary temperatures and offer a very convenient method of applying the alkali metals for chemical reactions. The amalgamating property of mercury is made use of in the recovery of gold and silver in certain mining processes.

**Salts of Mercury.**—Mercury is an example of an element which possesses a variable valence; that is to say, mercury salts exist in which the mercury ion has

valence of one ( $\overset{+}{\text{Hg}}$ ), there are also those salts in which the mercury ion has a valence of two ( $\overset{++}{\text{Hg}}$ ).

This property of variable valence is best illustrated by the chlorides. Mercury chloride ( $\text{HgCl}$ ) is calomel, a white, non-crystalline insoluble, non-poisonous, powder. Here one readily sees the mercury ion has in combination one ion of chlorine and is therefore monovalent. Mercury bichloride ( $\text{HgCl}_2$ ) is corrosive sublimate, a crystalline, soluble intensely poisonous solid. The mercury ion in this compound holds in combination two chlorine ions and is therefore bivalent. In compounds like  $\text{HgCl}$ ,  $\text{Hg}_2\text{O}$ ,  $\text{HgI}$ , etc., where the Hg ion is monovalent the mercury is said to be in the *mercurous* condition. The compounds mentioned then would be called mercurous chloride, mercurous oxide, and mercurous iodide. Such compounds as  $\text{HgCl}_2$ ,  $\text{HgO}$ ,  $\text{HgI}_2$  are called mercuric compounds (mercuric chloride, mercuric oxide, mercuric iodide).

The striking difference between mercurous and mercuric iodide is worthy of comment; mercurous iodide is bright yellow, mercuric iodide is scarlet red.

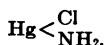
Both mercurous and mercuric compounds are employed extensively in medicine and surgery. The student should be thoroughly familiar with the two chlorides, especially since the substitution of one for the other is a very dangerous error.

On standing exposed to light, calomel becomes dark, owing to reduction and deposit of metallic mercury. Obviously at this stage it should not be administered.

Ammoniated mercury which is the principal constituent of Ammoniated Mercury Ointment is a mercuric compound in which one Cl is replaced by  $\text{NH}_2$ . Instead of



the formula then is



It is made by mixing a solution of bichloride and ammonia water when a white precipitate of ammoniated mercury is formed. When ammonia is added to a mercurous salt a black precipitate is formed. This is a simple method for ascertaining whether a given mercury compound is a mercurous or a mercuric compound.

#### SUMMARY OF CHAPTER XXII.

To the magnesium group belong magnesium, zinc, mercury, cadmium and glucinum. The most important member of the group for the nurse is magnesium, though mercury warrants description.

Magnesium occurs as the sulphate, the chloride, and the silicate. Epsom salt is the sulphate and asbestos is the silicate.

Metallic magnesium is light in weight and color. It burns in air with a brilliant, white light, forming  $\text{MgO}$ ; the substance when mixed with water is known as the milk of magnesia. Magnesium oxide in the stomach is converted into  $\text{MgCl}_2$ .

The spectroscope is an instrument for viewing the characteristic lines produced by molten metals or

their salts. By spectrum analysis elements may be detected in the sun.

The only metal liquid at ordinary temperature is mercury, which occurs free in nature and is recovered by distillation.

Amalgams are combinations of various metals like sodium, potassium, gold, and silver with mercury.

Mercury may be monovalent or bivalent. Calomel is the monochloride ( $\text{HgCl}$ ) in which the mercury is monovalent (in the mercurous state). Corrosive sublimate is mercuric chloride  $\text{HgCl}_2$  in which the mercury is bivalent. The differences between these salts are very important.

A convenient test for mercurous and mercuric salts is the addition of ammonia; mercurous salts are turned black; solution of mercuric salts yield a white precipitate on the addition of ammonia.

## CHAPTER XXIII.

### ALUMINUM—IRON—MANGANESE.

#### **ALUMINUM, Al. (At. wt. = 27).**

THE remarkably light metal aluminum is the element aluminum in the free state. This is an interesting metal from the stand-point of general knowledge. It suffices here to state that clay is an impure aluminum silicate which may vary in properties and uses according to its impurities. Aluminum hydroxide ( $\text{Al}(\text{OH})_3$ ) is used for clarifying solutions, and the class of compounds known as the alums are useful as styptics and astringents. The commonest representative of the alums is potassium alum (ordinary alum), which is a compound of potassium and aluminum sulphate ( $\text{Al} \cdot \text{K} \cdot (\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$ ) plus twelve parts of water of crystallization.

#### **IRON (FERRUM), Fe. (At. wt. = 56).**

Chemically and industrially iron is one of the most important elements. It plays an essential role in the living organism but a thorough knowledge of its salts is not required of the nurse. Only a few facts may be related.

The essential coloring matter of the red-blood cells is an iron compound. Iron is also found in bone.



Certain foods, like spinach and kale, are said to contain this element in larger amounts than other foods and, hence, are useful in furnishing a deficient organism with this essential element.

Iron, like mercury, has a variable valence, and exists in combination in the ferrous or ferric state. Thus  $\text{Fe}(\text{OH})_2$  is ferrous hydroxide,  $\text{FeCl}_2$  ferrous chloride; while  $\text{Fe}(\text{OH})_3$  is ferric hydroxide, and  $\text{FeCl}_3$  ferric chloride. The ferric salts are the more stable as the ferrous compounds are easily oxidized into the ferric variety. It has been found that the change from ferric to ferrous is a reduction process and, conversely, from ferrous to ferric an oxydative reaction; also the difference in the Fe ion is one of electrical charge.

There are about twenty-four official preparations of iron in the Pharmacopœia which may be said to be primary and eleven more made from these. The chemistry of these compounds can be readily understood from what has already been learned about chemical processes in general. One important preparation is the so-called *arsenic antidote* which is a mixture of magnesia and ferric hydroxide. The arsenic is precipitated by this mixture probably in the form of ferric and magnesium arsenite which is insoluble and can then be washed out of the stomach.

The preparation is made by mixing a solution of ferric sulphate (40 parts ferric sulphate in 125 parts of water) together with a thin paste consisting of 10 parts of magnesium oxide in 750 parts water. These two preparations should be kept made up separately ready for mixing.

**MANGANESE, Mn. (At. wt. = 55).**

Manganese dioxide,  $\text{MnO}_2$ , has already been referred to as an oxidizing agent. Manganese forms a great variety of compounds for the reason that it can exist in so many states of valence. It forms salts just as iron does and, on the other hand, enters into combinations to form negative ions. Potassium permanganate  $\text{KMnO}_4$  is an example of the latter. This compound loses its oxygen easily in either acid or alkaline solution. In contact with organic matter solutions of permanganate are reduced and lose the characteristic purple color, becoming a dirty brown. On account of its oxidizing power, permanganate is slightly antiseptic. It is used in medicine as an irrigant.

**SUMMARY OF CHAPTER XXIII.**

Aluminum is one of the lightest metals known. Its strength is out of proportion to its weight. The free element does not occur in nature but its silicate (clay) is abundant. It is separated from its compounds with great difficulty. However, electrochemistry has made it possible in recent years to produce aluminum at low cost.

The chief compounds of interest here are the hydroxide  $\text{Al}(\text{OH})_3$ , which is used for clarification purposes and the alums. Ordinary alum is a double sulphate of aluminum and potassium  $(\text{AlK}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O})$ , and is a representative of a type of double sulphates. There are also chromium alums and iron alums in which the

aluminum is replaced by chromium or iron as the case may be.

Iron is a very important element both from an industrial and a biological view point. The red coloring matter of the blood contains iron.

Iron occurs free in nature and is found extensively as the oxide.

Iron has a variable valence—it may be bivalent or trivalent.

The ferric compounds are more stable than the ferrous compounds.

Arsenic antidote is a mixture of magnesia and ferric hydroxide.

Manganese salts are very numerous because of the variable valence of the element.

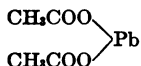
Manganese dioxide  $\text{MnO}_2$  and potassium permanganate  $\text{KMnO}_4$  are oxidizing agents. The latter in weak solution is an astringent.

## CHAPTER XXIV.

### LEAD—SILVER—PLATINUM.

**LEAD (PLUMBUM), Pb. (At. wt. = 207).**

THE heaviness of lead is proverbial though its specific gravity (11.4) is slightly less than mercury (13.9). Lead is bivalent and forms salts like  $\text{PbCl}_2$ . One of the most important salts in medicine is the acetate



the so-called *sugar of lead*, useful in skin diseases. Lead is slightly soluble in pure water and more soluble in water in which vegetation has fermented. Since even small amounts of lead taken into the body accumulate there and finally cause lead poisoning (plumbism) it is unsafe to use lead pipes for conducting drinking water.

**SILVER (ARGENTUM), Ag. (At. wt. = 108).**

Silver is moderately resistant to chemical action but is attacked readily by nitric acid which converts it into silver nitrate (lunar caustic  $\text{AgNO}_3$ ). Silver nitrate is reduced when it comes into contact with organic matter and metallic silver is deposited. This

is the reason that one's fingers are blackened by handling it. Silver nitrate forms compounds with albumins, *e. g.*, argyrol, which are used in the treatment of infections of mucous membranes.

The halogen compounds with silver, especially silver bromide, darken on exposure to light and for this reason are used to manufacture sensitized plates for photography. When light strikes these plates some slight change is produced according to the intensity and duration of the light. When this plate is put in some reducing agent like pyrogalllic acid ("developer") metallic silver is deposited where the light has affected the changes. This deposit of metallic silver will of course form shadows of varying degree. The unchanged silver bromide must be dissolved away before the other light comes into contact with the plate. This is done by washing in sodium hyposulphite,  $\text{Na}_2\text{S}_2\text{O}_3$  ("fixing" in "hypo"). The plate is washed again in water and dried. This "negative" is used to "print" the image on sensitized paper which is "developed" and "fixed" in the same manner as the plate.

**PLATINUM, Pt. (At. wt. = 195).**

Platinum on account of its usefulness and rare occurrence is worth more than gold. It is very highly resistant to chemical action and for this reason is very useful in chemical procedures. It may be heated in the air without being oxidized, hence the platinum needles with which the bacteriologist transfers cultures. On account of its high melting point and freedom from

oxidation, contact points for electrical apparatus are made of platinum. Points and knives for thermocautery are also made of platinum.

The action of platinum in a finely divided state in bringing about chemical reaction (catalyzer) has already been mentioned. Platinum chloride,  $\text{PtCl}_4$ , is used also as a catalyzer; for example, in the production of hydrogen by the action of  $\text{HCl}$  or  $\text{Zn}$ , a small amount of  $\text{PtCl}_4$  greatly accelerates the reaction.

**Colloids.**—Platinum is insoluble in water but if a strong current is allowed to pass between two platinum points in water, some of the metal goes into minute suspension in particles so small that they cannot be seen under a microscope. This is colloidal suspension. Only crystalloids (substances which can be crystallized) go into true solution. A colloidal suspension lies between a true solution and a fine suspension. Colloids are precipitated by boiling with acids. Other metals, as silver, gold, copper, etc., can be transformed into the colloidal state. These facts are related to give the student some sort of an idea of what is meant by the term *colloid* for many of the vital reactions of the cell life are now explained in terms of colloids.

#### SUMMARY OF CHAPTER XXIV.

Silver and lead are very similar chemically. Both are fairly resistant to chemical action; nitric acid, however, attacks both, forming nitrates.

Sugar of lead is lead acetate  $\text{Pb}(\text{OOC.CH}_3)_2$ . Lead is bivalent.

Metallic lead is slightly soluble in pure water and more soluble in slightly acidulated water. Lead poisoning may result from the consumption of water conducted in lead pipes.

Silver salts as well as lead are astringent. Silver nitrate and silver albuminates are used in the treatment of certain infections.

Silver salts find extensive use in photography, because of the fact that light darkens silver chloride, iodide, and bromide. Metallic silver is deposited in proportion to the intensity of the light. See text for description of process.

Platinum is very resistant to chemical action and for this reason is useful in many chemical processes. In the bacteriological laboratory the loops for transferring cultures are made of platinum because this metal can be heated to redness so many times without deterioration. Thermocautery points and contact points in electrical apparatus are made of platinum for the same reason.

Platinum salts (Pt. is tetravalent) like  $\text{PtCl}_4$  have the property of stimulating many chemical processes without entering into the final product (catalyzer). Opportunity is taken here to introduce the subject of colloids. An electrical current passed between two platinum points under water will project minute particles of platinum into fine (ultramicroscopic) suspension. Colloidal solutions are not true solutions, but stand somewhere between suspension and solution. The particles in colloidal suspension do not settle out on standing—and they cannot be seen with the aid of a

microscope. Colloids do not pass through animal membranes as solutions of crystalloids do. Colloids are precipitated by acids. Colloidal chemistry is coming into prominence in the study of pathological chemistry.



## CHAPTER XXV.

### CARBON, C.

(At. wt. = 12.)

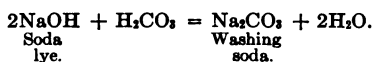
WE are familiar with the element carbon in non-crystalline forms as charcoal and graphite. One is surprised, however, to learn that the sparkling diamond is nothing more than pure crystallized carbon. The soft, friable, black substance seems to have nothing in common with the white, sparkling stone, the hardest substance known. On complete oxidation both substances yield carbon dioxide and nothing more. The French chemist, Moisson, was able to produce very small diamonds from charcoal, and more recently larger diamonds have been made.

**Distribution.**—When vegetable or animal material is heated in a closed vessel charcoal results, thus showing that carbon enters into the composition of these substances in relatively large amounts. We shall learn later that this element is the chief constituent of living matter. In combination with hydrogen, oxygen and nitrogen it is capable of forming an enormous variety of compounds. For example, the simple substance vinegar is composed of carbon, hydrogen and oxygen, and the highly complex and wonderful animal and vegetable cell substances are chiefly composed of these three elements plus nitrogen.

The vast oil and coal deposits are chiefly carbon and hydrogen. Carbon also occurs in minerals (carbonates).

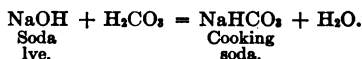
**Chemical Properties.**—Carbon is relatively inert chemically and combines with other elements only under the influence of heat. With lime, for example, under the influence of intense heat, the carbide of calcium  $\text{Ca}_2\text{C}$  is formed.<sup>1</sup> Heated in the air carbon is oxidized to form carbon dioxide,  $\text{CO}_2$ . This is the gas given off by animals during respiration and absorbed by the leaves of plants to be built up into complex vegetable matter. The ultimate product of oxidation of the vegetable and animal matter is  $\text{CO}_2$ . When carbon or any of the organic compounds are heated in an atmosphere poor in oxygen more or less of the monoxide ( $\text{CO}$ ) is formed. This gas is deadly to life. In the blood of animals it unites with the hemoglobin (red coloring matter) to form a stable compound and the animal becomes asphyxiated. Many cases of monoxide poisoning have been reported as resulting from the shutting up of stoves over night. Insufficient oxygen is supplied to the glowing carbon and carbon monoxide instead of carbon dioxide is formed. A "flare back" in a Bunsen burner brings about the same condition of insufficient oxygen supply and carbon monoxide is formed.

**Carbonates.**—Carbon dioxide in water solution forms the unstable acid  $\text{H}_2\text{CO}_3$ , carbonic acid. In the presence of hydroxides of metals the carbonate is formed. For example:



<sup>1</sup> When water is added to  $\text{Ca}_2\text{C}$  acetylene (which see) is formed.

When only half the quantity of the hydroxide is present, a hydrogen or acid carbonate is formed.



The bicarbonate (acid carbonate) of sodium is cooking soda to which reference has already been made.

The carbonates of magnesium, of calcium, of barium and of strontium are insoluble, therefore the addition of a soluble carbonate to a solution of any salt of the above elements brings about a precipitate. This property is made use of in the separation or estimation of these elements.

#### SUMMARY OF CHAPTER XXV.

Charcoal, graphite, and diamond are forms of the element carbon.

Carbon is found abundantly in nature: as carbonates it occurs in mineral deposits, and the element enters into a large number of compounds with H, O and N to form the chief constituents of animal and vegetable matter.

Carbon is relatively inert chemically. It combines, however, with various elements on heating. Carbon is tetravalent. Completely oxidized it forms,  $\text{CO}_2$ , which is given off in expiration. Incomplete oxidation results in the formation of a highly poisonous compound, carbon monoxide (CO), which forms a stable union with the red-blood cells.

## CHAPTER XXVI.

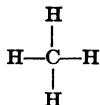
### COMPOUNDS OF CARBON WITH HYDROGEN.

THE study of the various compounds of carbon with hydrogen and with hydrogen and oxygen is known as *organic* chemistry because the living cells and their products are such compounds of carbon.

**Marsh Gas.**—The simplest of the organic compounds is the gas which bubbles up from the stagnant water overlying decomposing vegetable matter. If this gas be collected and mixed with the proper proportion of air it forms an explosive mixture showing that it is an easily oxidizable compound. This is the gas which often collects in mines and explodes when a miner's lighted torch is brought in.

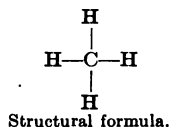
By analysis we learn that marsh gas is composed of four parts of hydrogen to one part of carbon. The analyses of other carbon compounds show that carbon combines with the equivalent of four hydrogen atoms. For example, in carbon dioxide, since oxygen has twice the valence of hydrogen, carbon is said to have a valence of four. One atom of carbon will hold four atoms of chlorine in combination,  $\text{CCl}_4$ . Since the valence of chlorine is one (remember  $\text{HCl}$ ) thus carbon is again shown to be tetravalent. Many other examples can be cited.

Marsh gas is written



which is our nearest approach to the representation of our conception of the relation of the atoms to each other. In reality we believe that the hydrogen atoms are arranged at the points of a tetrahedron with the carbon atom as the centre. When we increase the number of carbon atoms in a compound one can readily see that the position occupied by the atoms may make a considerable difference in the character of the compound. A formula showing the relative positions of the atoms is said to be the *structural* representation, while the formula indicating simply the relative number of atoms, *e. g.*,  $\text{CH}_4$ , is called the *empirical* formula.

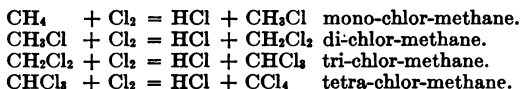
$\text{CH}_4$   
Empirical formula.



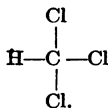
It will be seen later that two compounds may possess entirely different properties, but have the same empirical formula: only by the structural formula could one distinguish one from another when they are referred to.

**Methane.**—Organic chemistry treats of the various compounds built up on the basis of marsh gas as a unit. The chemical name of marsh gas is methane. A mixture of methane and chlorine in diffused day-

light will react to form chlorine substitution products of methane.



**Chloroform.**—The third product tri-chlor-methane is the compound familiar to us as chloroform. The structural formula is



Chloroform is a heavy, mobile liquid, having a characteristic odor and sweet taste. It boils at 62° C., and the vapors are not inflammable as in the case of ether, though when the vapors are brought in contact with a flame they are slightly oxidized, forming carbonyl chloride, a dangerous gas. For this reason care should be exercised in the use of chloroform as an anesthetic near a lamp or gas flame. Commercially, chloroform is produced by the action of bleaching powder on alcohol. *Iodoform* is tri-iodo-methane, that is, iodine is substituted for three hydrogen atoms in methane.



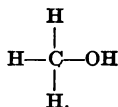
Iodoform is a yellow, crystalline solid, with a characteristic penetrating odor. Its antiseptic properties are due to the slow liberation of free iodine.

**Methyl Alcohol.**—If we treat mono-brom-methane with silver oxide in the presence of water we obtain methyl hydrate and silver bromide.



Methyl hydrate is the compound which we obtain in the destructive distillation of wood and call wood alcohol. Just as the hydroxyl group OH is characteristic of a base when joined to a metal, it is when joined to an organic group the characteristic of an alcohol. An alcohol then consists of a hydroxyl group joined to an organic group (radical).

Methyl alcohol is represented structurally by the following formula:



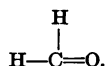
If R represent any organic radical then an alcohol may be represented by R.OH.

Methyl alcohol is lighter than water, and mixes with it in all proportions. This alcohol is used as a solvent in industrial processes and the fumes often cause blindness in the workmen. When taken internally it is a poison and may cause death. On account of its cheapness there is a temptation to use it as an adulterant in the cheaper wines and whiskies which is of course illegal.

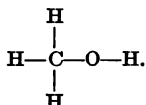
**Formaldehyde.**—The disinfectant used for fumigating rooms where patients with contagious diseases have been is formaldehyde gas. Formalin is a 40 per cent. solution of the gas in water. An exceedingly small

amount is able to inhibit the growth of bacteria and larger amounts kill them. It is poisonous for animals when taken internally. Strong solutions burn the skin, and there are those who have an idiosyncrasy for it to the extent that even very weak solutions cause violent skin reactions.

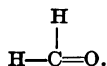
The chemical constitution of formaldehyde is simple: its structural formula is



We see that this compound is methane in which two H atoms are replaced by O. Observe that the valence of carbon is four and that they are satisfied. In methyl alcohol the O has one valence bound by H, viz.:



but in formaldehyde both bonds of the oxygen are attached to the C. Also it will be seen that formaldehyde has two atoms of hydrogen less. We, therefore, see the relation between methyl alcohol and formaldehyde. The taking away of hydrogen means oxidation, that is, some oxygen has combined with these two hydrogen atoms to form water which splits off, leaving



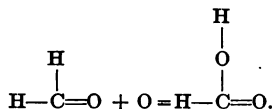
The group





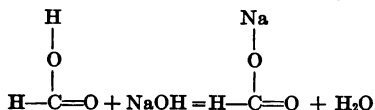
(written also CHO) is called the aldehyde group. It is characteristic of the aldehydes because no compound is an aldehyde unless it possesses such a group, and all compounds possessing this group are aldehydes. The aldehydes are made by oxidizing alcohols. In the example here given, methyl alcohol vapor passed over heated copper or platinum wire in the presence of air is oxidized to formaldehyde. If R represent any organic group (radical) then an aldehyde of this radical may be represented by: R.CHO.

**Organic Acids.**—If formaldehyde is treated with an oxidizing agent like potassium permanganate  $\text{KMnO}_4$  it is oxidized to formic acid:

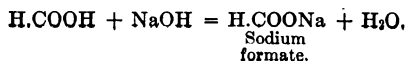


Formic acid is a corrosive, colorless compound, with a penetrating odor occurring in the bodies of ants (Latin *formica* = an ant). It also occurs in the hairs of certain caterpillars and in the stings of nettles.

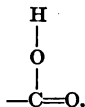
It is important to observe that only one hydrogen is replaceable by the metal when the acid is neutralized by a base. The equation for the reaction of sodium hydroxide on formic acid is represented by the following:



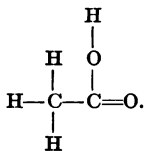
This equation is usually written



Formic acid is the simplest of the organic acids. The carboxyl group



(written usually COOH) is characteristic of organic acids. If R represent any organic radical then an organic acid may be represented by R.COOH. In the case of formic acid R is only H, but in acetic acid (vinegar), for example, R is a methyl group, CH<sub>3</sub>. Now substituting CH<sub>3</sub> for R in our general formula for an organic acid R.COOH, we have CH<sub>3</sub>.COOH (acetic acid). The structural formula of acetic acid may be represented as follows:



### SUMMARY OF CHAPTER XXVI.

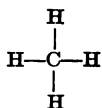
Organic chemistry deals with the compounds of C, H, O, N. The natural compounds made up of these elements are the result of plant or animal life so that the term organic is applied to all such compounds to distinguish them from the mineral or inorganic substances.

Marsh gas is the simplest of the organic compounds, and the nucleus about which the more complicated are arranged. It is very important to learn the various

## 148 COMPOUNDS OF CARBON WITH HYDROGEN

substitution products of methane, for this is the foundation of organic chemistry.

Carbon is tetravalent. The H atoms of marsh gas ( $\text{CH}_4$ ) are supposed to be arranged in space about the C atom as a centre. A convenient hypothesis is that each H atom is placed at the angle of a tetrahedron (a four-sided solid). It is conceivable that the figure is equilateral if the four valences of the carbon are satisfied by the same atoms. Such a formula the representation on paper is thus:



is called the structural formula, while  $\text{CH}_4$  is the empirical formula. Two substances may have the same empirical formula but different structural formulas.

The four H atoms of methane can be replaced by other monovalent atoms. For example, chloroform is tri-chlor-methane (three H atoms replaced by chlorine). Chloroform is a colorless, mobile liquid. Iodoform, a yellow, crystalline solid, is an analogous compound; instead of three chlorine atoms there are three iodine atoms substituted for three hydrogen atoms.

If a hydrogen is replaced by a hydroxyl group the result is methyl alcohol,  $\text{CH}_3\text{OH}$ .

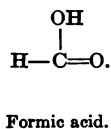
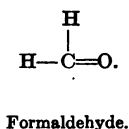
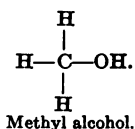
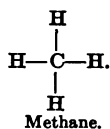
An alcohol consists of a hydroxyl group joined to an organic radical,  $\text{R}-\text{OH}$ .

If two hydrogen atoms are replaced by an oxygen atom the result is formaldehyde,  $\text{HCHO}$ .

The characteristic of the aldehyde group is  $\text{—CHO}$ . Any aldehyde may be represented thus:  $\text{R—CHO}$ .

If two hydrogens are replaced by an atom of oxygen, and one hydrogen replaced by a hydroxyl group the result is an organic acid, formic acid,  $\text{H.COOH}$ . The  $\text{—COOH}$  group is characteristic of an organic acid,  $\text{R—COOH}$ .

An acid is an oxidized aldehyde, and aldehyde is an oxidized alcohol, and an alcohol is an oxidized hydrocarbon.



## CHAPTER XXVII.

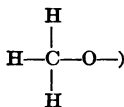
### ETHERS.

If methyl (wood) alcohol  $\text{CH}_3\text{OH}$  is allowed to come in contact with metallic sodium, the alcoholate of sodium is formed:

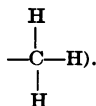


We have already seen that sodium has a strong attraction for the halogens (iodine, chlorine, bromine and fluorine), forming with them salts or halides. We have also seen that from methane,  $\text{CH}_4$ , a compound of chlorine, bromine or iodine may be formed  $\text{CH}_3\text{Cl}$ ,  $\text{CH}_3\text{Br}$ ,  $\text{CH}_3\text{I}$ . Then if we put either of these three substances in solution with a sodium compound there would be a strong tendency to form  $\text{NaCl}$ ,  $\text{NaBr}$  or  $\text{NaI}$ , according to the methyl halide present. If  $\text{CH}_3\text{O.Na}$  be brought in contact with  $\text{CH}_3\text{I}$ , then we would have  $\text{NaI}$  formed leaving two compounds with unsatisfied or unsaturated bonds:

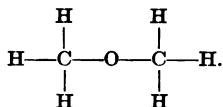
$\text{CH}_3\text{O}$ —(or



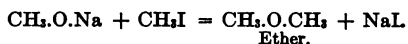
and  $\text{CH}_3$ — (or



The natural result is the joining of these unsaturated bonds thus:



As a matter of fact this actually happens and we have therefore the equation:



This compound  $\text{CH}_3\text{O.CH}_3$ , written also



is called an ether because of its low boiling-point and elastic property. This (methyl-methyl-ether) is the simplest of the ethers. Later we shall learn that the ether used as an anesthetic (ethyl ether) is of the same type of compound, namely, two organic radicals (R) joined with oxygen:



It is not essential that the two radicals are the same: the compound

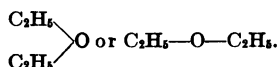


( $\text{R}_1$  representing any other radical, for example  $\text{C}_2\text{H}_5$ ), is still an ether.

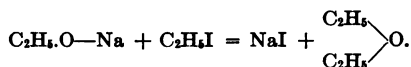
## SUMMARY OF CHAPTER XXVII.

Ethers have the constitution  $R-O-R_1$ , in which  $R$  and  $R_1$ , represent any organic radical as  $CH_3$ ,  $C_2H_5$ ,  $C_3H_7$ , etc. It is not necessary that the two organic groups have the same constitution.

The ether given for anesthesia is ethyl ethyl ether,



Ethyl ethyl ether can be made by the reaction of sodium ethyl alcoholate and ethyl iodide:



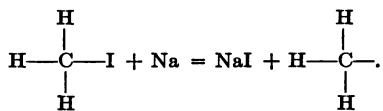
In practice this substance is produced by the action of sulphuric acid on ethyl alcohol. On distillation ether comes over.

Ether is a colorless, mobile liquid, having a characteristic odor, and a burning, sweet taste. Its specific gravity is about 0.7. It is highly inflammable. It boils at  $36^\circ$  C. Ether for anesthesia contains about 4 per cent. ethyl alcohol and a small amount of water.

## CHAPTER XXVIII.

### THE MARSH GAS SERIES.

It has been shown that marsh gas  $\text{CH}_4$  in the presence of chlorine, iodine or bromine in the sunlight will gradually form methyl chloride, iodide or bromide. If we heat one of these compounds with sodium, remembering the great affinity sodium has for the halogens (Cl, I, Br), we would expect the sodium to combine with the halogen to form a salt. Suppose we take  $\text{CH}_3\text{I}$  as an example, because this compound is more easily handled than  $\text{CH}_3\text{Cl}$  or  $\text{CH}_3\text{Br}$ , on account of the lower boiling-point of the former. If metallic sodium is brought into contact with  $\text{CH}_3\text{I}$  in ether solution (no water must be present because the sodium would combine with the water to form  $\text{NaOH}$ ) sodium iodide is formed.



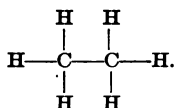
This leaves an unsaturated compound  $\text{CH}_3-$  (that is, a carbon atom with only three bonds satisfied). According to chemical laws this compound readily combines with an available body. What, then, is the most available body for this unsaturated compound



to attach itself to? The molecules of substances are so minute that the smallest amount we can appreciate must contain millions of molecules. When the reaction takes place between one molecule of each, it suffices to write the equation of single molecules. In the reaction we have millions of molecules of unsaturated compounds and as we would expect, they pair off, combining with each other after the manner of the following:

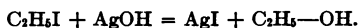


The structural formula of  $\text{CH}_3\text{—CH}_3$  is



This substance is called ethane and is generally written  $\text{C}_2\text{H}_6$ . It is a gas similar to methane and found in petroleum and in the neighborhood of oil wells. Chemically it reacts like methane. For example, with iodine each of the H's in turn are replaceable, forming ethyl mono-, di-, tri-, etc., iodide ( $\text{C}_2\text{H}_5\text{I}$ ,  $\text{C}_2\text{H}_4\text{I}_2$ ,  $\text{C}_2\text{H}_3\text{I}_3$ ,  $\text{C}_2\text{H}_2\text{I}_4$ ,  $\text{C}_2\text{HI}_5$ ,  $\text{C}_2\text{I}_6$ ). We come to know  $\text{C}_2\text{H}_5$  as an organic radical, ethyl.

**Alcohol.**—Then if we treat  $\text{C}_2\text{H}_5\text{I}$  with silver hydroxide we expect to get an alcohol (see methyl (wood) alcohol), according to this equation:



This is ethyl alcohol, the ordinary alcohol we know in medicine.

**Properties.**—Pure alcohol is a colorless, volatile liquid, having an agreeable odor and burning taste. It mixes readily with water and ether in all proportions. It is lighter than water, boils at a lower temperature ( $78^{\circ}$  C.) and freezes at a much lower temperature ( $-130^{\circ}$  C.). Alcohol is easily oxidized: fine platinum wire accelerates oxidation of the fumes and will thus set fire to alcohol vapors. It burns in air with a non-luminous, sootless flame. Alcohol is useful on account of its solvent power for oils, resins, and alkaloids.

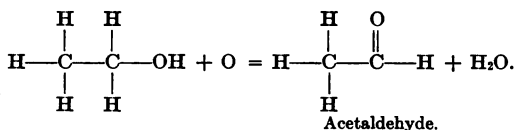
**Source of Alcohol.**—Yeast is a unicellular organism belonging to the fungi. Cultures viewed through a microscope are found to consist of large numbers of single cells. These cells are able to convert certain sugars into carbon dioxide and alcohol. This property is taken advantage of for the commercial production of alcohol. The simplest example is the manufacture of wine in which the natural grape sugar (glucose or dextrose) is changed by the yeast to alcohol and carbon dioxide. The carbon dioxide coming off as a gas gives the appearance of boiling so that the process is called fermentation (fervere=to boil). Wine contains from 10 to 20 per cent. alcohol. In the distillation more alcohol than water goes over on account of the lower boiling-point of the former and the product known as brandy contains about 50 per cent. alcohol. Brandy, of course, contains higher alcohols (fusel oils), which give it the peculiar taste. If lime is added and another distillation carried out the product is almost pure alcohol. Any vegetable containing starch may be used to manufacture alcohol, but the starch must first be

broken down into sugars before the yeast can utilize it. To accomplish this step a ferment is obtained from sprouting barley and this (diastase) added to the cooked starch in the process known as *malting*. Malt sugar, grape sugar, and dextrins result. Yeast is added and alcoholic fermentation begins. If hops and malt are fermented beer is made. Corn, rye, and potatoes are also used, but here the mash is distilled and whisky is the result. Beer contains from 5 to 8 per cent. alcohol and whisky contains about the same amount as brandy (40 to 50 per cent.).

**Fermentations in General.**—Fermentation is now applied generally to mean the changes brought about by the class of substances known as ferments. The chemical composition of these substances is unknown: they are products of living cells and act according to certain laws. If they enter into combination they are immediately set free, for a small amount of ferment is capable of changing large amounts of substances if given sufficient time.

Strictly speaking, fermentation means the destruction of sugars with the production of carbon dioxide and alcohol or organic acids, in contrast to *putrefaction*, which is the decomposition of proteins with the production of ammonia and foul-smelling gases. Where microorganisms are capable of inducing fermentation or putrefaction, the former takes precedence over the latter. This means that in a decomposing mixture, as a rule, protein decomposition does not take place and foul odors do not arise until all the sugars are destroyed by fermentation.

**Acetaldehyde.**—We remember that formaldehyde  $\text{HCHO}$  was formed by the *oxidation* of methyl alcohol,  $\text{CH}_3\text{OH}$ , or the *reduction* of formic acid,  $\text{HCOOH}$ . So acetaldehyde is formed by the oxidation of ethyl alcohol  $\text{C}_2\text{H}_5\text{OH}$ .



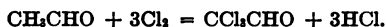
Acetaldehyde is written  $\text{CH}_3\text{CHO}$ .

Acetaldehyde is a volatile, colorless liquid with a suffocating odor. It is little used in medicine but its polymer,<sup>1</sup> paraldehyde, is a very useful and safe hypnotic.

**Paraldehyde.**—When a drop of sulphuric acid is added to acetaldehyde a condensation occurs. Three molecules of acetaldehyde combine with one another to form one molecule of paraldehyde  $(\text{CH}_3\text{CHO})_3$ , a volatile liquid with a pungent taste capable of producing sleep with very little depression of the heart or ill after-effects. It should be kept in a cool place.

**Chloral.**—This well-known hypnotic is a chlorinated acetaldehyde.

On treating acetaldehyde with dry chlorine gas the following reaction takes place:

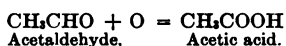


Chloral is then tri-chlor-acetaldehyde. Chloral, itself, is a colorless, oily liquid, but on the addition

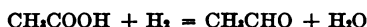
<sup>1</sup> Polymer is a molecule consisting of two or more molecules condensed into one. The verb is polymerize.

of water, crystals of chloral hydrate form,  $\text{CCl}_3\text{CHO} \cdot \text{H}_2\text{O}$ . This is the compound usually prescribed.

**Acetic Acid.**—The next step in the oxidation of ethyl alcohol, after acetaldehyde is formed, is the corresponding acid (compare formic acid).



and the reverse is true,



by the reduction of acetic acid we obtain acetaldehyde.

Acetic acid is the chief constituent of vinegar. The usual method of manufacture is by growing a special fungus or mould in weak alcohol. When apple juice is used fermentation first takes place, and the sugar is changed to alcohol (hard cider) then acidification begins. Vinegar contains from 1 to 3 per cent. acetic acid. The slimy sediment sometimes seen consists of masses of the mould which forms the vinegar (mother of vinegar).

#### SUMMARY OF CHAPTER XXVIII.

The marsh gas series consists of carbon-hydrogen compounds of gradually increasing complexity, beginning with methane  $\text{CH}_4$  and progressing by the successive additions of  $\text{CH}_2$ . Ethane is  $\text{CH}_4 + \text{CH}_2$  or  $\text{C}_2\text{H}_6$ . Propane =  $\text{C}_2\text{H}_6 + \text{CH}_2 = \text{C}_3\text{H}_8$ , etc.

Ethane may be formed from methane by first producing the mono iodo methane  $\text{CH}_3\text{I}$  and treating this

compound with metallic sodium in a water-free medium. The marsh gas series may be built up in this manner.

The alcohols may be formed by treating the alkyl iodide with  $\text{AgOH}$ .

Ethyl alcohol  $\text{C}_2\text{H}_5\text{OH}$  is the common (grain) alcohol of commerce. It is a colorless, volatile liquid, lighter than water, miscible in all proportions with water and ether. The commercial source is the fermentation of sugars by yeast.

Ferments are substances of unknown composition which are capable of bringing about repeated chemical changes without entering into the end-products. A small amount of ferment can change an indefinite amount of sugar if given time enough.

The aldehydes of the marsh gas series are made by oxidizing the corresponding alcohol. Acetaldehyde is the result of the partial oxidation of ethyl alcohol. Its formula is  $\text{CH}_3\text{CHO}$ .

Paraldehyde is formed by the condensation of three molecules of acetaldehyde.

Chloral is a chlorinated acetaldehyde;  $\text{CCl}_3\text{CHO}$ .

Acetic acid (vinegar) may be produced by further oxidation of ethyl alcohol or acetaldehyde. It is manufactured by growing a special fungus (mother of vinegar) in weak alcohol. Vinegar contains 1 to 3 per cent. acetic acid. The formula for acetic acid is  $\text{CH}_3\text{COOH}$ .

Let the student compare the formulas of ethane, ethyl alcohol, acetaldehyde, and acetic acid.

## CHAPTER XXIX.

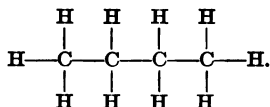
### THE PARAFFINS.

PETROLEUM is a mixture of a large number of compounds composed of carbon and hydrogen. The simplest of these products, methane (marsh gas), we have already described. This has the composition  $\text{CH}_4$ , and the next member of the series is ethane which we learned is  $\text{C}_2\text{H}_6$ . It will be seen that the difference between these members is  $\text{CH}_2$ . Ethane was made from methane by first preparing the iodide  $\text{CH}_3\text{I}$  and treating this with sodium. By following this method the various members beginning with methane may be prepared:

Methane	$\text{CH}_4$	gas
Ethane	$\text{C}_2\text{H}_6$	"
Propane	$\text{C}_3\text{H}_8$	"
Butane	$\text{C}_4\text{H}_{10}$	liquid
Pentane	$\text{C}_5\text{H}_{12}$	"

and so on.

The structural formula for butane as an example is:



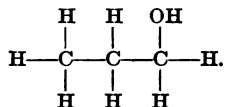
Observe that each end carbon has three hydrogen atoms, while the included carbons hold only two.

In building these compounds a methyl group,  $\text{CH}_3$ , is added, but it will be remembered that one of the end hydrogens is split off to make room for the attachment of the carbon atom. This explains why the difference between successive members is  $\text{CH}_2$  rather than  $\text{CH}_3$ .

As we proceed upward in this series the compounds become less volatile, that is, their boiling-point increases. These compounds have been isolated from natural petroleum, but it is a less difficult task to build them from methane and ethane than to attempt to separate them. Gasoline is a mixture of several members of this series.

**Derivatives of the Hydrocarbons.**—Since methane and ethane furnished us on oxidation the corresponding alcohols (methyl and ethyl) it is possible to prepare the alcohols of the other members of this series. Hence we have propyl alcohol, butyl alcohol, and so on. We also have the corresponding aldehydes, ethers, and acids.

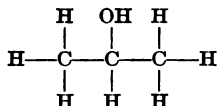
**The Higher Alcohols.**—Ethane we have seen is  $\text{C}_2\text{H}_6$  and ethyl alcohol is  $\text{C}_2\text{H}_5\text{OH}$ . Then if propane is  $\text{C}_3\text{H}_8$ , propyl alcohol is  $\text{C}_3\text{H}_7\text{OH}$  and its structural formula is:



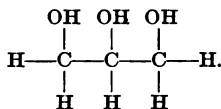
This we call a primary alcohol, but if the hydroxyl group, which we remember is the characteristic group



of an alcohol, is attached to the middle C as in the following:



we name it a *secondary* alcohol. We shall find little need of distinguishing between primary and secondary alcohols, but we shall find it to our advantage to pay close attention to the following compound:



Observe that there are three alcohol groups here, and we call it a tri-hydroxy alcohol or tri-atomic alcohol. This is the mon-atomic alcohol already described with the further replacement of two hydrogen atoms by hydroxyl groups. If there were two hydroxyl groups it would be called a di-atomic alcohol.

**Glycerine.**—The tri-atomic alcohol, whose formula has just been stated, is the substance known in medicine as glycerine.

**Properties.**—Glycerine is so named on account of its sweet taste. As we ordinarily know it, this tri-hydroxy alcohol is an odorless, clear, thick liquid, although in its pure state it is crystalline. It has the property of absorbing water (hygroscopic), and on this account a thin film will keep surfaces moist. It is useful as a vehicle in pharmacy.

Glycerine is made from fats by treating them with

superheated steam, and it is a by-product in the manufacture of soap. The reactions will be discussed under fats.

### SUMMARY OF CHAPTER XXIX.

The paraffin series consists of *saturated* hydrocarbons of the marsh gas series. Various members of the series are found in petroleum. The simplest member is methane,  $\text{CH}_4$ . Ethane is  $\text{C}_2\text{H}_6$ . By the addition of a  $\text{CH}_2$  group to any member the next higher is obtained. The first four members are gases—the remainder are liquids (excepting the most complex members which are solid at ordinary temperature).

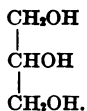
The structural formulas of these compounds are very important. The empirical formulas tell only a very small part of the story of their composition.

Derivatives of these hydrocarbons are known. These correspond to the various compounds of methane formed by replacing the H atoms with other elements or groups. It will be seen that more derivatives of the same kind are obviously possible; for example, while in the case of methane only one alcohol was possible, in propane three are possible. Likewise more chlorine derivatives are possible in the higher members of the series.

In a primary alcohol the hydroxyl group is attached to the end carbon; in a secondary alcohol this group is attached to the second carbon, etc.

A monatomic alcohol possesses one hydroxyl group: a diatomic alcohol, two groups, etc.

Glycerine is a common example of a tri-atomic alcohol. It has the composition  $(\text{CH}_2)_2 \cdot \text{CH} \cdot (\text{OH})_3$ .

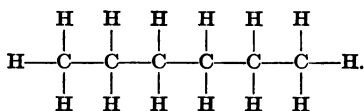


It is so named on account of its sweet taste. It is a thick, clear syrup, miscible in water in all proportions and crystallizable in the absolutely pure state. Glycerine is a by-product of soap manufacture.

## CHAPTER XXX.

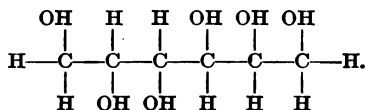
### SUGARS.

THE compound  $C_6H_{14}$  is one of the paraffins, that is, it is composed of six methyl groups connected to form a chain in the following manner:



Obviously several alcohol derivatives of this compound are possible. If one of the end H's were replaced by OH the compound would be hexyl alcohol,  $C_6H_{13}OH$ . The other extreme, or complete hydroxylation, would of course consist of one OH group for every C atom. There could not be any more because if more than one OH group attaches itself to a C atom two of them combine and are split off as water ( $H_2O$ ).

The compound resulting from the replacement of one of the H's on each C atom then is hexatomic alcohol:



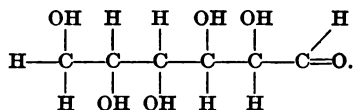
(Notice that the hydroxyl groups are not all on same side.)

This compound is found in nature and is called *mannite*. It is a white crystalline substance with a sweet taste. It is fermentable and in other ways it is similar to sugars.

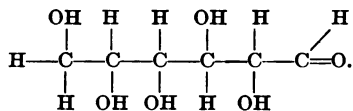
Mannite is used extensively in bacteriology to differentiate certain kinds of bacilli, especially the organisms causing bacillary dysentery. Just one step, one slight chemical change, brings us to *mannose*, which is a sugar. By oxidizing one of the end groups we obtain an aldehyde. (Remember that formaldehyde was produced by oxidizing methyl alcohol.) Mannose is a hexatomic alcohol (mannite) with an aldehyde



group on the end:



If now we change the relative positions of the H and OH groups, we have the formula:



This is grape sugar (glucose).

**Properties of Grape Sugar.**—Grape sugar is commonly known as glucose. (*Glyc* or *gluc* stem in a word indicates sweetness, and the suffix *ose* shows that the sub-

stance is a sugar in the chemical sense.) Glucose is a white crystalline substance, soluble in water and slightly soluble in alcohol. It possesses a pleasant, sweet taste, but no odor. Glucose occurs in nature in combinations known as glucosides. It occurs in grapes uncombined, hence the name grape sugar. Commercially it is obtained by boiling starch with dilute acid (sulphuric acid is generally used because it can be so easily eliminated afterward). The reason why glucose can be obtained from starch in this manner will appear later.

**The Polariscope.**—A ray of light is said to be a vibration in ether. If we liken it to the vibration of a harp string we find that the excursions are not all in the same place. In other words, if the string were stretched in the direction exactly perpendicular to the earth the vibrations when the string is struck would not be confined to excursions north and south or east and west, but would swing to any or all points of the compass. Such is our conception of the vibrations of a ray of light. It is possible to keep the vibrations in the same plane: suppose two plane boards were placed one on each side of the string in such a manner that the string could move to and fro in one direction. The string would vibrate then in one plane. The same thing can be accomplished with light by allowing the ray to pass through a prism. The light waves vibrate in one plane just as the string did and the result is *polarized* light. If a ray of polarized light is passed through solutions of sugars, the vibration plane is turned to the right or left. Suppose the

plane of vibration of a particular beam of polarized light is exactly vertical. The beam is passed through a solution of glucose. Now the plane of vibration is not vertical but has been rotated to the right (clockwise) so many degrees. Glucose then is said to be dextro-rotary (dextra = right). The amount of rotation is proportionate to the amount of sugar present. It is therefore possible to determine how much glucose is present in a solution without going through the process of recrystallizing it several times to purify it and finally drying and weighing it. The instrument with which one measures the rotation of the plane of light is called a *polariscope*. The number of degrees a prism must be rotated to bring the plane back to the original position is read and from this the amount of sugar can be calculated. The polariscope is used extensively in the sugar industries and in analyses of sugar-containing substances. It is essential in accurate analysis of urines which contain sugar.

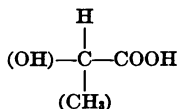
**The Asymmetric Carbon Atom.**—The power to rotate the plane of polarized light is due to the presence of a carbon atom which has all four of its bonds satisfied by different kinds of *groups*. Such a carbon atom is said to be asymmetric. If any two of these groups are the same, the carbon atom is not asymmetric.<sup>1</sup>

Suppose, for example, that we substitute for three of the hydrogen atoms in methane the following groups ( $\text{CH}_3$ ), ( $\text{OH}$ ), and ( $\text{COOH}$ ), *i. e.*, a methyl, a hydroxyl, and a carboxyl group. One H atom remains.

<sup>1</sup> Asymmetric means *without symmetry* (not symmetrical).

We have now a compound containing an asymmetric carbon atom, and according to the above statement should rotate the plane of polarized light.

This compound

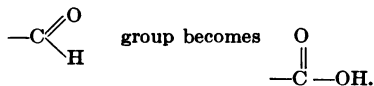


is lactic acid. It rotates the plane of polarized light to the *right*. Should we change the relative positions of the methyl and carboxyl groups we still have lactic acid, but this lactic acid rotates the plane to the *left*. This is the basis upon which our theories of optical activity of chemicals is built.

**Reducing Power of Sugars.**—Returning to the graphic formula of dextrose we find that it contains an aldehyde group,



We remember that one of the chief characteristics of an aldehyde is its power to reduce substances, that is, take oxygen from other substances and become oxidized itself. To form an acid; the



We therefore suspect dextrose of being able to reduce substances. Testing the substance for this property we find that it is so. If an alkaline solution of a copper



salt be boiled in the presence of dextrose the copper is reduced and settles out as a bright red, fine precipitate (copper oxide,  $\text{Cu}_2\text{O}$ ).

The solution best adapted for applying this fact in testing for dextrose is composed of copper tartrate and sodium hydroxide. This solution is usually known as Fehling's solution.

**Levulose.**—Another sugar, which on analysis would yield carbon, hydrogen and oxygen in exactly the same proportions as found in dextrose, rotates the plane of light to the left (contra clock-wise) and is, therefore, called levulose. It is the sugar found in honey and also in fruits. The graphic formula of levulose differs from that of dextrose in that the CHO group is not at the end of the chain but inside. It is therefore not an aldehyde strictly but a ketone on account of the relative position of the aldehyde group. The strictly aldehyde sugars are called *aldoses* and ketone sugars named *ketoses*. Levulose reduces Fehling's solution like dextrose and can be fermented (*i. e.*, broken up by ferments to form acids and carbon dioxide), though less readily than dextrose.

**Monosaccharids.**—Dextrose and levulose are typical examples of the several simple *hexoses* or monosaccharids. The group which they represent are called monosaccharids to distinguish them from the group, the members of which are formed by the chemical union of two simple sugars.

## SUMMARY OF CHAPTER XXX.

Hexane has the formula  $C_6H_{14}$ . It is possible to replace six H atoms with hydroxyl groups. No more OH groups can be joined to this compound because if more than one such group is attached to the same carbon atom, water would be split off by the union of the two.

One hexa-hydroxy-hexane found in nature is mannite. This alcohol is very similar in its physical properties to sugars. It is used in bacteriological work.

If one of the end alcohol groups of mannite be oxidized to an aldehyde group, the result is mannose, a sugar.

By shifting the relative positions of the middle H and OH groups in mannose, grape sugar, another sugar is obtained. Glucose is a white crystalline substance, soluble in water and slightly soluble in alcohol, with a sweet agreeable taste and without odor. It is found in grapes and other fruits. It can be obtained by hydrolyzing starch.

The polariscope is an instrument by means of which the rotation of the plane of polarized light is determined. In passing through sugar solutions the plane of polarized light is rotated to the right or left, and the amount of rotation is directly proportional to the amount of sugar present. This is a simple means for the determination of sugars in solution. The power to rotate the plane of light depends upon the presence of an asymmetric carbon atom in the compound.

On account of the presence of an aldehyde group

in the hexoses (six carbon sugars), these compounds are capable of reducing substances like copper tartrate in alkaline solution. Both aldoses and ketoses (which see) have reducing power. The copper solution in alkaline tartrate used for testing sugars is commonly known as Fehling's solution.

Levulose is fruit sugar. Solutions of this sugar rotate the plane of polarized light to the left or in opposite direction the rotation by grape sugar (dextrose). Levulose is a ketose. It is found in honey.

Dextrose and levulose are examples of simple sugars (hexoses) and are classed as monosaccharids in order to distinguish them from the di- and polysaccharids formed by the union of two or more monosaccharids.

There are also sugars containing only five carbon atoms termed *pentoses*. They are sometimes found in human urine but are of more interest to the physiological chemist.

## CHAPTER XXXI.

### POLYSACCHARIDS.

**Cane Sugar.**—The familiar crystalline substance used extensively for sweetening is a result of chemical union between the two monosaccharids already described. Dextrose and levulose combined with the loss of one molecule of water, form cane sugar or saccharose. Both dextrose and levulose have the empirical formula  $C_6H_{12}O_6$ . Combination of these two produces  $C_{12}H_{24}O_{12}$ , but there is lost in the process of uniting one molecule of water  $H_2O$ , therefore we have as the empirical formula of saccharose  $C_{12}H_{22}O_{11}$ .

**Properties of Cane Sugar.**—Cane sugar is found in the juice of sugar cane, in beets, in bananas and other fruits. It crystallizes easily from concentrated solutions, and is therefore easily obtained in a pure state. Solutions of cane sugar rotate the plane of polarized light to the right. Fehling's solution is not reduced, showing that the aldehyde groups of the simple sugars composing it are completely masked.

What evidence is there that cane sugar is composed of dextrose and levulose? Boil some cane sugar with weak hydrochloric acid (1 part concentrated HCl to 100 c.c. solution of cane sugar) for two hours. The hydrochloric acid may be removed by precipitating

with lead or silver salts. Filtering through charcoal gives us a clear, colorless solution, which we may now compare with part of the original solution (cane sugar solution). The original solution rotated, the polarized light to the right, now the solution is levorotary (rotating to the left). The original solution did not reduce Fehling's solution but the new clear liquid does so very vigorously. From the original solution only one substance could be crystallized, from the recent liquid two substances can be separated, the one dextrorotary the other levorotary. Further chemical test show these substances to be dextrose and levulose. It is thus proved that hydrochloric acid adds the lost molecule of water and splits cane sugar into its simple sugars. The word hydrolysis has been proposed to describe this process (hydro = water; lysis = breaking down). Thus this word is applied to all processes which add water chemically to a substance and divide it into its component parts.

**Invertase.**—We found that boiling with hydrochloric acid changed the power of the solution to rotate the plane of light, that is, it reversed the direction of rotation to almost an equal extent in the other direction. It practically *inverted* the rotation—so we speak of the products of hydrolysis as *invert* sugar, and of the process as an *inversion*. Living yeast cells or extracts of yeast cells are also capable of bringing about the inversion of cane sugar, acting at the temperature of the human body. The substance in the yeast which acts in this manner is capable of converting many times its own weight of cane sugar,

it obeys certain laws of rate of reaction, and it is killed by heat. A substance answering such a description is known as a ferment. Names of ferments of this class end in *ase*, and the stem of the word indicates either their action or the substance they change. The logical name for the inverting ferment is *invertase*. Invertases occur in the intestinal juices having been secreted by the mucous lining of the intestine. It is by means of these ferments that the higher sugars (disaccharids) are reduced to simple sugars for absorption in the process of digestion. Invertases are found in small amounts in the blood.

#### OTHER DISACCHARIDS.

**Milk Sugar.**—If some copper sulphate is added to skim milk and then a small amount of sodium hydroxide, all the white substance of the milk will be coagulated. Should we filter the mixture we obtain a clear liquid from which on evaporation a white, sweet substance crystallizes. This is milk sugar. The Latin word for milk is *lactis*, the ending *-ose* indicates a sugar; milk sugar is therefore well named *lactose*.

Solutions of lactose rotate the plane of polarized light to the right, and also reduce Fehling's solution. Lactose is sweet, but not so sweet as glucose (dextrose), and it is not so soluble in water as the latter, nor is it so easily fermentable.

Boiling with acid (hydrolysis) breaks up lactose into dextrose and another simple sugar, galactose. The mucous lining of the intestines yields a ferment known as lactase, capable of hydrolyzing lactose.

In the digestion of lactose, this lactase hydrolyzes it before it is absorbed by the villi of the intestinal wall.

Lactose is unique in being the only disaccharid of animal origin. It is found in the milk of all mammals: about 5 per cent. of cow's milk and about 7 per cent. of human milk is lactose. From whatever source it is obtained it is the same chemically.

**Malt Sugar.**—Malt sugar is a disaccharid composed of two molecules of dextrose, yielding these on hydrolysis by boiling with an acid or under the influence of maltase. Maltose is sweet, is dextrorotary, and reduces Fehling's solution. It is found in malt where it is produced from starch by the action of a ferment derived from germinating (sprouting) grain.

### STARCHES.

The reserve material of plants is stored in the form of pure white insoluble substances called starches. Most of the dry material of the potato, for example, is starch; corn also contains a high percentage of this carbohydrate. A prominent characteristic of starch is its reaction with iodine to form a blue color, which disappears on heating but reappears on cooling. To test any mixture for the presence of either of these substances (starch or iodine) one has merely to add the other.

Under the microscope starch is found to be composed of small, oval grains, having on them lines forming elongated ellipses drawn about a common centre. Starches from different sources have different

shapes and markings, but so far as is known they are the same chemically. In fact the exact composition of starch is not known. It is evident from many studies that it is composed of a group of molecules of dextrose, but how many and their manner of combination has not been determined. Boiling with mineral acids decomposes starch with the formation of dextrose (glucose) and dextrans. A ferment, amylase (amylum = starch), found in the saliva, in pancreatic juice and in sprouting grain is also capable of hydrolyzing starch in the same manner.

The starch grains have an outer coat of cellulose (wood tissue) which protects it from ferment action until it is broken. Heat ruptures this coat so that cooking prepares the starch for digestion. In this we find the reason for the indigestibility of raw starch, except by cows and goats, whose intestinal juices contain ferments capable of dissolving cellulose.

Starch is the great source of commercial glucose. Grain alcohol is made by hydrolysis and fermentation of starches. (See Ethyl Alcohol.)

**Cellulose.**—Cellulose is a complex compound of sugars and starches in chemical combination. It is characterized by its extreme insolubility. Wood fibre is largely cellulose.

#### SUMMARY OF CHAPTER XXXI.

Polysaccharids are carbohydrates made up of two or more monosaccharids. Disaccharids are polysaccharids consisting of two monosaccharids.



Cane sugar is a disaccharid composed of dextrose and levulose; that is, 1 molecule of cane sugar is composed of 1 molecule of dextrose, plus 1 molecule levulose, minus 1 molecule water. Dextrose ( $C_6H_{12}O_6$ ) + levulose ( $C_6H_{12}O_6$ ) =  $C_{12}H_{24}O_{12}$ , but one molecule  $H_2O$  is lost, so the empirical formula of cane sugar is  $C_{12}H_{22}O_{11}$ .

Cane sugar is also called saccharose (sweet sugar) or sucrose.

Cane sugar does not reduce Fehling's solution. We know that dextrose which is included in cane sugar reduces Fehling's (evidence of an aldehyde group). We therefore conclude that the aldehyde group in cane sugar is masked.

Boiling with acids adds a molecule of water to cane sugar and splits it into dextrose and levulose.

Invertase is a ferment found in yeast which is capable of hydrolyzing cane sugar into the same products as acids do.

Milk sugar (lactose) is a disaccharid occurring in milk. It is composed of two monosaccharids: dextrose and galactose. Lactose is the animal sugar—it has not been found in plants. It is dextrorotary. Lactase is the ferment which hydrolyzes lactose. Lactose reduces Fehling's solution.

Maltose is malt sugar, composed of two molecules of dextrose. It is dextrorotary and reduces Fehling's solution. Maltase is the ferment which hydrolyzes malt sugar.

Starches are the reserve materials of plants. They are composed of an unknown number of molecules of

dextrose. Starch grains from different sources have different sized grains and diverse markings on them.

Starch grains have an outer covering of cellulose (wood fibre); until this covering is broken starch grains are not soluble in water. Boiling ruptures the capsule and a part of the contents is soluble in water. Boiling with dilute acids hydrolyzes starch into dextrose and dextrans. Diastase, a ferment found in the saliva, and amylase, a ferment found in the intestine, also hydrolyze starch.

Starch reacts characteristically with iodine to form a blue color.

Inulin is similar to starch, but it is a polysaccharid composed of an unknown number of molecules of levulose.

Cellulose is also a polysaccharid, characterized by its extreme insolubility. Wood fibre is chiefly cellulose.

## CHAPTER XXXII.

### THE DIGESTION OF CARBOHYDRATES.

CARBOHYDRATE is a comprehensive term, including simple and complex sugars and starches. Generally speaking, the carbohydrates are so named because they contain carbon in combination with hydrogen and oxygen, the latter two in the same ratio as they occur in water (2H to 1O).

**In the Mouth.**—Starches are the only carbohydrates which undergo chemical change in the mouth. Cooked starches are hydrolyzed by the amylase (diastase) in the saliva to dextrose and dextrans. Raw starch is indigestible. The action of the saliva is stopped as soon as the food comes into contact with the hydrochloric acid of the stomach, for the starch splitting ferment is active only in alkaline medium.

**In the Stomach.**—Practically no carbohydrate digestion takes place in the stomach. The food entering the stomach is already mixed with saliva which continues to act until the acid of the stomach stops it. There is no absorption of carbohydrates from the stomach.

**In the Intestines.**—As soon as the food reaches the small intestine the reaction is again changed to alkaline. The intestinal juices coming from the pancreas and from the wall of the intestine contain ferments which

split starches and disaccharids. All carbohydrates capable of digestion are reduced to monosaccharids before being absorbed; for example, starches and maltose are reduced to dextrose; saccharose (cane sugar) to dextrose and levulose, and lactose (milk sugar) to dextrose and galactose. Dextrose (grape sugar), levulose (fruit sugar), and galactose (from milk sugar) are absorbed as such without change.

**Fate of the Sugars.**—Some of the sugars are destroyed by fermenting bacteria in the intestine before they can be absorbed. Most of them are absorbed and either stored in the body as glycogen or fat or oxidized to form heat. On oxidation they yield  $\text{CO}_2$  and  $\text{H}_2\text{O}$ . If too large amounts are absorbed, the part which the body cannot care for is excreted in the urine.

**Energy Furnished by Carbohydrates.**—The body is capable of oxidizing carbohydrates to their ultimate products ( $\text{CO}_2$  and  $\text{H}_2\text{O}$ ), therefore they furnish as much heat to the body as they would in bomb calorimetric combustion. While the various sugars differ slightly in the number of Calories per gram they yield, the average for a normal diet is 4 Calories per gram.

**Storage of Carbohydrates.**—When the body requires less carbohydrates than is assimilated the excess of simple sugars is used to build up a compound corresponding to vegetable starch. This substance of unknown composition, called glycogen, is formed in the liver and stored there or in muscles until it is needed. When dextrose is needed the glycogen is broken down to yield it. Glycogen gives a red color with iodine instead of the blue as in the case of starch.

Sugars may also be converted into fats in the body and stored as such.

The average normal dietary for a 150 pound man, at moderate work, contains about 500 grams (dried) of mixed carbohydrates per day.

#### SUMMARY OF CHAPTER XXXII.

The digestion of carbohydrates which takes place—

In the mouth: Starches hydrolyzed by amylase to dextrins and dextrose.

In the stomach: No carbohydrate digestion after the contents become acid. No absorption.

In the intestines: Fresh amylase from pancreas and wall of intestine continue hydrolysis. All digestible polysaccharids reduced to monosaccharid, and absorbed as such.

Fate of absorbed monosaccharids: (1) either oxidized to give heat to body, or (2) changed into glycogen (animal starch) and stored in liver and muscles, or (3) changed into fats and deposited in body, or (4) in case of extra large amounts which body cannot handle the excess is excreted in the urine.

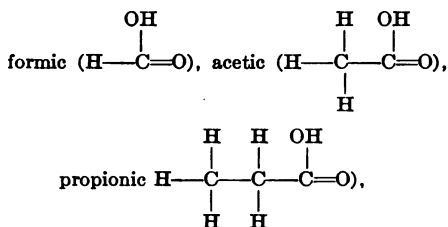
One gram of carbohydrate furnishes about 4 Calories of energy to the body.

The average moderate dietary contains about 500 grams (water free) carbohydrate per day.

## CHAPTER XXXIII.

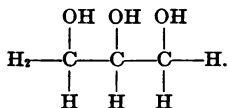
### FATS.

It will be remembered that the organic acids:



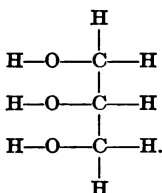
etc., were described as being the first three of a series of acids increasing in complexity. This series is the fatty acid series, so named because these acids occur in animal and vegetable fats in combination with other substances.

In the discussion of alcohols, glycerine was mentioned as being a tri-atomic alcohol (having three hydroxyl groups):

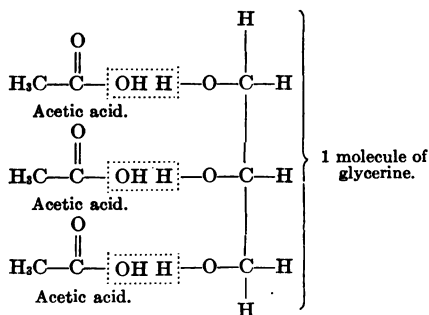


Now we are allowed to arrange this molecule in space in any position we choose so long as the relative

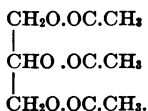
position of one group to another remains the same. Thus we may write glycerine as follows:



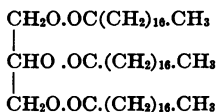
Suppose now we place an acid with a carboxyl (COOH) group near each hydroxyl group. It will be remembered that when two hydroxyl groups approach one another in a chemical reaction they unite and split off water, leaving the groups to which they were attached united by the free bonds thus liberated.



Here three molecules of acetic acid (vinegar) combine with one molecule of glycerine. Three parts  $\text{H}_2\text{O}$  are split off and a simple *fat* is formed. The formula for this fat is



We have seen, then, that a fat is a chemical combination of glycerine (glycerol) with three molecules of some fatty acid. Observing the formula just given, and remembering the graphic formula of the paraffin series, we realize that the three  $\text{CH}_3$  groups may be supplemented by successive additions of  $\text{CH}_2$  groups. In other words we can place as many  $\text{CH}_2$  groups between  $\text{HOOC}$ — and  $-\text{CH}_3$  as we wish. Accomplishing this, suppose we stop at sixteen groups, thus arriving at the formula:



which happens to be the composition of one of the body fats *stearin*. Stearic acid is  $\text{CH}_3(\text{CH}_2)_{16}.\text{COOH}$ .

### THE BODY FATS.

The animal fats consist largely of mixtures of three different fats: *stearin*, *palmitin*, and *olein*.

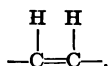
**Stearin.**—*Stearin* is the hardest and most insoluble of the body fats. It is practically insoluble in water but goes into solution in alcohol and ether. It is the fat which makes tallow solid. Vegetable fats also contain it.

**Palmitin.**—*Palmitin* is also a solid fat and is present in animal fats in larger amounts than *stearin*. There are two  $\text{CH}_2$  groups less in palmitic acid than in stearic, that is, its formula is  $\text{CH}_3(\text{CH}_2)_{14}.\text{COOH}$ .

**Olein.**—*Olein* occurs in greater quantities in vegetable fats than in animal fats, although it is present



in all animal fats. At ordinary temperature it is a colorless oil, lighter than water. It is not exactly like stearin and palmitin in composition, for some of the carbon atoms have only one H and are, hence, bound to the next C, by *two* bonds, thus



We call oleic acid an *unsaturated* acid because it takes up chlorine, bromine, iodine, and other elements so easily.

#### VEGETABLE FATS.

Olive oil, cottonseed, peanut, and corn oil, etc., are mixtures of various fats, hence their relation to the animal body is essentially the same as that of animal fats.

#### SOAPS.

If fats are boiled with lye there is evidently a chemical change, for the two substances lose their identity and two new substances appear, namely, soap and glycerine. On analysis we find that the soap is a combination of the alkali (sodium) with the fatty acid forming a salt. For example, the stearin would form sodium stearate  $\text{CH}_3(\text{CH}_2)_{16}.\text{COONa}$ . The OH of the lye (NaOH) went to replace that lost in the glycerine when it was joined to the fatty acid to make the fat. Any of the metals capable of combining to form salts can be combined with a fatty acid to form a soap. Most of the soaps, except sodium and potassium, are insoluble in water and, hence, useless for cleansing,

although they may be used medicinally as ointments. The process of breaking up a fat to form a soap, setting free glycerine is known as saponification (saponis, Latin = soap; facio = to make).

Soaps clean greasy substances by emulsifying the fats so that water will remove them.

### THE DIGESTION OF FATS.

The digestion of fats does not begin until the food has passed from the stomach into the intestine.

Here fat digestion proceeds in two different steps: (1) the pancreatic juice and the bile emulsify the fats; (2) an enzyme (lipase) of the pancreatic juice hydrolyzes the fats into glycerine and fatty acids.

These are absorbed by certain cells in the mucous membrane of the intestine, and it is thought they are built up to form body fats before being taken into the blood stream.

The newly formed body fats are stored in various parts of the body until needed for energy production.

### FOOD VALUE OF FATS.

Fats are burned in the body, furnishing heat and forming  $\text{CO}_2$  and  $\text{H}_2\text{O}$ . Having less oxygen in the molecule in proportion to carbon and hydrogen than is found in carbohydrates, they require more oxygen from the air for the same amount of  $\text{CO}_2$  given off. Therefore, by determining the amount of oxygen consumed, the amount of  $\text{CO}_2$  and heat given off, and the loss in weight, we are able to ascertain whether

carbohydrates or fats are being oxidized to furnish body heat.

Calorimetric experiments have shown that 1 gram of fat yields to the human body about 9.3 Calories. An average liberal dietary for a man contains about 50 grams of fat per day.

### SUMMARY OF CHAPTER XXXIII.

Fats are hydrolyzed by superheated steam, forming glycerine and an organic acid. One molecule of glycerine (a tri-atomic alcohol) is combined with three molecules of some fatty acid to form fat. In the process three molecules of water are lost. In the hydrolysis these three molecules of water must be supplied.

The character of the fat, olein, palmitin, stearin, etc., depends upon the fatty acid constituent. The glycerine is the same in all fats.

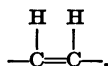
The essential part of vinegar, acetic acid, is the lowest of the series of fatty acids. It has the composition  $\text{CH}_3\text{COOH}$ . Other members of the series are obtained by adding  $\text{CH}_2$  or multiples of this methylene group, *e. g.*,  $\text{CH}_3\text{CH}_2\text{COOH}$ ,  $\text{CH}_3(\text{CH}_2)_2\text{COOH}$ , etc. Stearic acid contains sixteen of these groups:  $\text{CH}_3(\text{CH}_2)_{16}\text{COOH}$ .

The body fats are mixtures of stearin, palmitin and olein. Stearin is the hardest and most insoluble of the body fats. It has the highest melting-point.

Palmitin is also a solid fat. Most of the body fat consists of palmitin.

Olein is fluid at ordinary temperatures. It occurs

in larger quantities in vegetable fats than in animal fats. Oleic acid differs from stearic and oleic acid in its inner structure. There are not so many H atoms in proportion to the number of C atoms, for some of the carbon atoms have only one H attached and consequently they are attached to one another by a double bond thus:



This kind of an acid is said to be *unsaturated* because it easily absorbs iodine, bromine, chlorine, hydrogen, etc.

Vegetable fats are essentially the same as animal fats. They contain greater proportions of olein and are therefore more fluid. It is apparent that any desired consistency of fats may be attained by varying the proportions of stearin (hard fat) and olein (fluid fat).

If fats are boiled with any alkali, hydrolysis takes place: glycerine is set free and the fatty acid combines with the alkali metal to form *soap*. For example, stearin boiled with lye forms glycerine plus sodium stearate (soap).

A soap is the fatty acid salt of an alkali. Sodium and potassium soaps are soluble. Magnesium and calcium soaps are insoluble. It is now plain that *hard* water is uneconomical for cleaning purposes. Soaps emulsify fats and allow them to be removed in suspension.

Fats are digested in the small intestine. The bile

and pancreatic juice emulsify them and lipase, the fat splitting enzyme, hydrolyzes them into glycerine and fatty acid. These two substances pass through the wall of the intestine and are reformed into fats. Fats are either burned to furnish energy or stored until needed. Blood serum taken soon after meals is often turbid on account of the fat suspended in it.

Each gram of fat furnishes about 9.3 Calories to the body on oxidation.

## CHAPTER XXXIV.

### BENZENE SERIES.

#### COAL TAR.

WHEN coal is heated to a high temperature in a vessel from which all the air is excluded, three classes of substances are formed: (1) substances which are gaseous at ordinary temperature called illuminating gas; (2) substances gaseous at high temperatures, but liquid at ordinary temperatures called coal tar; and (3) is the solid residue coke.

It is the second class, coal tar, which merits our interest. From it are obtained the raw materials for many well-known substances. Many drugs like acetanilid, phenacetin, etc., aniline dyes and perfumes and other complex derivatives are obtained from this black, tarry, viscous, repelling liquid.

If coal tar is distilled and the condensed products collected in water a separation can be made into two classes of substances: those lighter than water and those heavier than water. Of course the lighter oils rise to the surface of the water where they may be collected. One notices the odor of carbolic acid and finds that on washing with lye that this odor is removed and that the liquid is clearer and has a more pungent odor. By distilling this carefully and collecting the

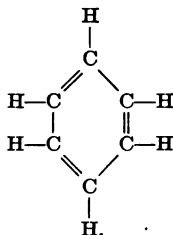
portions which come over at certain temperatures, the liquid may be divided into several fractions which possess different boiling-points and specific gravities, as well as differ in their chemical constitution.

### BENZENE.

The lowest member of the series is commonly known as benzene (not the benzine which is obtained from petroleum). It is a colorless liquid, having a characteristic odor, boils at  $80^{\circ}$ , and burns with a luminous flame.

**Chemical Properties.**—By analysis it is found that carbon and hydrogen are present in the ratio of 1 to 1, that is, one atom of carbon to one atom of hydrogen. The formula then may be CH or any multiple of this. The molecular weight determination tells us that the molecule weighs about 103 times as much as hydrogen, therefore the formula must contain six CH groups. ( $C = 12$ ,  $H = 1.08$ ,  $CH = 13.08$ .  $103 \div 13.08 =$  about 6.) Then the formula must be  $(CH)_6$  or  $C_6H_6$ . According to our conception of the combinations of carbon and hydrogen in chain formation, there ought to be more hydrogens present to satisfy the C atoms. A  $C_6$  chain would contain 14 H's. Therefore, some of the carbon atoms are not satisfied by H atoms, and since all the carbon atoms seem to be satisfied (benzene is not very active chemically), they must be joined to one another in certain places. The valence of carbon is four. Taking all these facts into consideration

with many others, a different arrangement of the C and H atoms have been arrived at, namely, a ring formula:



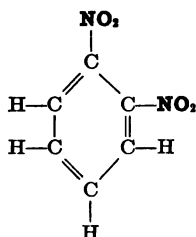
This is known as the benzene ring.

It is without the sphere of present considerations to discuss the methods of experimentation and reasoning by which this formula was deduced. It suffices to say that it offers the best explanation for the phenomena observed in work with this very interesting compound. It may be worth while to mention the effect of the halogens on benzene. In the sunlight a mixture of benzene and bromine unite to form a series of brom-benzenes, mono-, di-, tri, etc., up to six. They have the formulas  $C_6H_5Br$ ,  $C_6H_4Br_2$ ,  $C_6H_3Br_3$ , etc., from which it will be seen that in each instance the bromine replaces the H atoms. The same thing is true when the various nitrates of benzene are prepared.

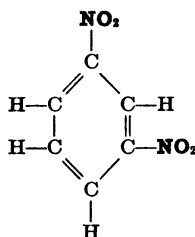
Does it matter which H is replaced? Any number of monobrom- or mono-nitro-benzenes have been prepared and in all instances they were identical, showing that all the H atoms have the same relation to the ring. If now we form di-brom- or di-nitro-ben-



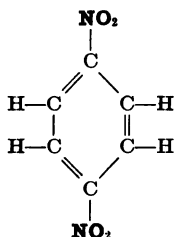
zenes we begin to observe differences in them. For example:



has different physical properties from



and still another di-nitro-benzene has different properties from either:

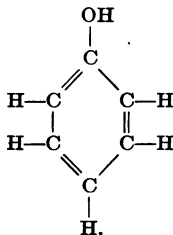


thus showing that the relative positions of the groups to one another make differences in the chemical com-

pounds. These positions have names: numbering the top C as 1 and proceeding clockwise, we observe that in the first compound, the hydrogen attached to 1 and 2 were replaced (this is called *ortho*-di-nitro-benzene, or the two nitro groups are said to be *ortho* to one another). In the second example 1 and 3 were replaced, thus producing a *meta* compound, and in the third example 1 and 4 were replaced, making *para*-di-nitro-benzene.

Benzene is one of the most important compounds because it is the foundation of a large series of highly interesting and important substances found in medicine and physiology.

**Substitution Products of Benzene.—Phenol.**—It was stated that the lighter oils from which benzene was prepared had the odor of carbolic acid. It has been found that carbolic acid does occur in both the first and the second separations and can be recovered in a pure crystalline state. Enquiring into the chemical constitution of carbolic acid it has been found that it is a benzene in which one of the H atoms has been replaced by a hydroxyl group, having the formula  $C_6H_5-OH$  or



From our knowledge of the methane series it would appear that carbolic acid could be made from benzene. We can prepare mono-brom-benzene in the method already given. Boil this with silver oxide and carbolic acid is formed (compare preparation of methyl alcohol from methane, page 144). From the resulting mixture carbolic acid can be crystallized.

Thus carbolic acid is not an acid at all. Instead of an organic acid group (COOH) it contains a hydroxyl group (OH) making it an alcohol. It should not be called carbolic acid: phenol is the proper name and suitable because it signifies that a phenyl group ( $C_6H_5$ ) is joined to an alcohol group (*ol* signifies an alcohol).

*Properties*—Phenol crystallizes in beautiful, clear, colorless needles, which develop a brownish-red color on exposure to light over a long period. It is soluble in alcohol and ether. Saturation in water yields a 5 per cent. solution.

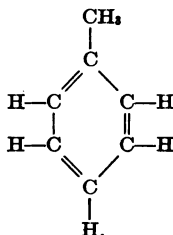
Phenol is a protoplasmic poison and for that reason is used as an antiseptic and germicide. It is interesting to note that phenol was the first antiseptic used in surgery. Lister, the founder of antiseptic surgery, used it in such great quantities that he himself and his assistants developed phenol poisoning.

Caustic soda unites with phenol to form sodium phenolate,  $C_6H_5O.Na$ .<sup>1</sup> Similar compounds are formed by potassium, barium, etc. Sulphates form a loose compound with phenol, and ethyl alcohol neutralizes

<sup>1</sup> Sodium phenolate is not hydrolyzed by water into NaOH and phenol, as we would expect if phenol corresponded exactly to an ordinary alcoholate. The presence of the phenyl group then confers some sort of acidic properties to the alcohol group.

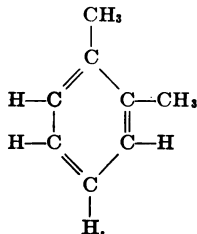
its poisonous effects. Therefore, alcohol or any soluble, non-poisonous sulphate forms the logical antidote in cases of poisoning by phenol.

**Toluene.**—It ought to be possible to substitute a methyl group for any of the H atoms of benzene. This has been done, and the natural product having identical composition is also found in the light oil obtained from the coal tar. This is toluene and has the formula:



It is slightly heavier than benzene and not so volatile. It is added in experiments in physiological chemistry to prevent bacterial growth. It inhibits ferment action to a slight extent.

**Xylene.**—The next product in the benzene series is xylene, which is di-methyl benzene; that is, two of the hydrogen atoms are replaced by methyl groups. One formula is



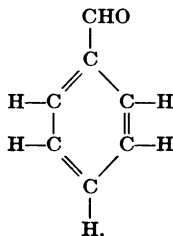
Of course there are two other possibilities as in the case of the nitro-benzenes. Xylene is heavier than toluene and boils at a higher temperature. It is used as a dehydrating agent in treating pathological sections and also to clean cedar oil from lenses. Xylene dissolves paraffin very easily.

**Other Members of the Benzene Series** differ in the number of hydrogen atoms replaced by methyl groups. They are of interest to the organic chemist.

**Cresols.**—The alcohol of toluene corresponding to phenol (the alcohol of benzene) is known as cresol. Several of these alcohols are possible according to the position of the alcohol group. The cresols are antiseptic and find use in practical disinfection.

#### ALDEHYDES OF THE BENZENE SERIES.

**Benzaldehyde.**—Just as we have aldehydes of the methane series, so there are aldehydes of the benzene series. Remembering that the group —CHO is characteristic of the aldehydes, benzaldehyde would be

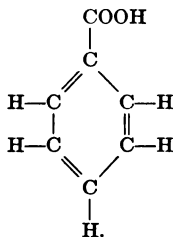


The relation to benzene is apparent. Of course the aldehyde can not be formed directly from the benzene

by oxidation because another C atom is necessary. It is therefore necessary to use the methyl substitution product (toluene) for oxidizing, to obtain benzaldehyde. This is the substance lending the peculiar, pungent odor to bitter almonds. Benzaldehyde is, therefore, called "oil of bitter almonds."

**Benzoic Acid.**—Much interest has been attached in late years to benzoic acid on account of its use as a preservative in certain foodstuffs. It may well be called a coal-tar derivative, because it can be obtained by oxidizing toluene. It is found in gum benzoin, and certain balsams. It crystallizes in white needles or plates.

The constitution of benzoic acid is:



Taken into the body it is excreted in the urine in combination with amino-acetic acid, forming hippuric acid. The latter occurs normally in the urine of horses and cows.

#### SUMMARY OF CHAPTER XXXIV.

Coal tar results from the air-free distillation of coal. Coal tar contains a large number of substances which

form the basis of the manufacture of drugs, dyes, perfumes, and other chemicals.

Coal tar may be redistilled and by condensing it into water two portions are collected: one lighter than water, the other heavier. Both of these portions may be fractionally distilled and various compounds separated.

Benzene is a light, colorless liquid found in the lighter portion. By analysis benzene has been found to be composed of ( $C_1H_1$ ). Molecular weight determinations show the relative weight of the molecule is 103. Then mol. wt. of CH (13.08) divided into 103 gives about 6. Therefore, the formula of benzene =  $(CH)_6$  or  $C_6H_6$ . The generally accepted structural formula is the ring of Kekulé. In the ring formula each H atom bears the same relation to the molecule. The H atoms may be replaced by halogens forming compounds like  $C_6H_5Cl$ ,  $C_6H_4Cl_2$ , etc., up to  $CCl_6$ .

The H atoms may be replaced by  $NO_2$  groups to the same extent. All mono-nitro benzenes possess the same characteristics, but there are three different di-nitro-benzenes. Nitro groups replacing H atoms next to one another form *ortho*-nitro-benzene; if one H lies between them a *meta* compound results, and if two H atoms lie between them a *para* compound results. Ortho, meta, and para refer to position. These names are not restricted to nitro compound; they apply to other substitution products as well.

Phenol is hydroxy benzene  $C_6H_5OH$ . It is really an alcohol and not an acid, though phenol seems to possess more affinity for bases than ordinary alcohols (see note

page 196). Carbolic acid is an erroneous name for phenol. Phenol can be prepared from benzene by the same process in which methyl alcohol is produced from methane.

Toluene is methyl benzene, *i. e.*, one of the H atoms is replaced by a methyl group ( $\text{CH}_3$ ).

Xylene is di-methyl-benzene. Two H atoms are replaced by methyl groups. Here again three xylenes are possible—ortho, meta, and para.

Benzaldehyde is commonly known as oil of bitter almonds. It is benzene in which an H atom has been replaced by an aldehyde group,  $-\text{CHO}$ .

Benzoic acid occurs in gum benzoin and balsams. It can be made by oxidizing toluene or benzaldehyde. Instead of an aldehyde group as in benzaldehyde there is an acid group (carboxyl,  $\text{COOH}$ ) replacing an H atom.



## CHAPTER XXXV.

### NITROGEN, N.

(At. wt. = 14.)

NITROGEN is an element of great importance. It is involved in all branches of chemistry.

**Occurrence.**—Nitrogen is present to some extent in all living matter: it is one of the principal elements in the living cell. Approximately four-fifths of the air is free, inert nitrogen.

**Properties.**—Nitrogen belongs in the same group with arsenic, antimony, phosphorus, and bismuth. It is a tasteless, colorless, odorless gas characterized by its extreme inertness, even at high temperatures. The nitrogen of the air is so inactive chemically that it is only with great difficulty that it can be made to combine with other elements. This has been accomplished, however, by means of electric sparks. Certain bacteria which grow on the roots of clover, peas, etc., seem to fix atmospheric nitrogen with such ease that they serve to fertilize the soil. Such crops are often planted in rotation and cultures of nitrogen-fixing organisms sown on the soil.

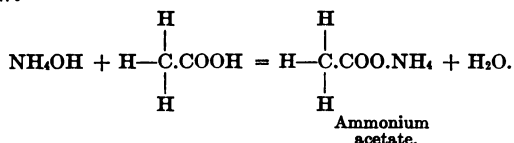
**Compounds of Nitrogen.**—Nitrogen occurs in the air as  $N_2$ . On passing an electric spark through a mixture of nitrogen and hydrogen, ammonia is formed. This has the composition  $NH_3$ , showing us that nitrogen is trivalent, although at times it is pentavalent.

**Properties of Ammonia.**—Ammonia dissolves in water, forming ammonia water ( $\text{NH}_4\text{OH}$ ) or ammonium hydroxide. In this state it is active chemically like an alkali, the group  $\text{NH}_4$  corresponding to a metal united with a hydroxyl group. Ammonia is an irritating, penetrating gas, readily detected even in small quantities in the air. The effect on the skin and eyes is such that strong solutions should be handled with care.

Ammonia can be liquefied by pressure. In returning to the gaseous state it absorbs large amounts of heat. This fact is utilized in modern ice-making and refrigerating machines. The ammonia is liquefied by compression and cooled by tap water. It is then allowed to expand into large pipes coiled in brine. The brine is thus cooled and circulated by pumps.

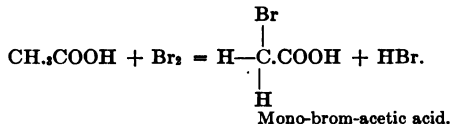
**Salts of Ammonia.**—Ammonium hydroxide neutralizes acids with the production of salts, as for example, ammonium chloride, ammonium sulphate, ammonium acetate, etc. These salts are, as a rule, very soluble in water, are white crystalloids and are volatile. In medicine they are administered as heart stimulants and expectorants. There are certain double salts of interest to the analytical and research chemist.

The reaction between acetic acid and ammonia will interest us on account of its bearing on the chapter to follow.

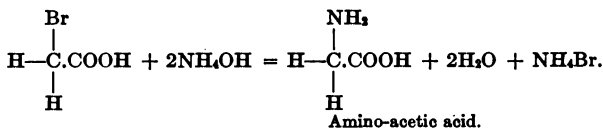


Notice the position of the ammonium group.

**Amino-acids.**—By treating acetic acid with bromine the hydrogens of the methyl group ( $\text{CH}_3$ ) are successively replaced by bromine. The first reaction is as follows:



If, now, we treat one molecule of this substance with two molecules of ammonia we find that the ammonia will combine with the bromine to form ammonium bromide, leaving an unsaturated bond which will take up the other molecule of ammonia.



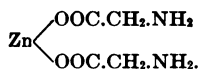
This substitution product in a fatty acid is of very great importance in physiological chemistry because of the relation of these acids to proteins (the chief constituents of living cells).

Amino-acetic acid is the representative of a class. Going back to the study of fatty acids it will readily occur to one that there are many such compounds possible. For instance, the fatty acids higher in the scale of complexity (containing more methyl groups), as propionic, butyric, etc., acids may contain amino groups in the place of hydrogens of the methyl groups. There are other possibilities: we have learned that

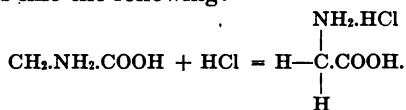
the relative positions of groups make considerable differences in the chemical as well as physical properties of substances. Then the amino group ( $\text{NH}_2$ ) may replace different hydrogen atoms. Also, the occurrence of two or three or four amino groups in higher fatty acids gives us the corresponding di, tri, etc., amino-acids.

**Properties.**—Amino-acids are crystallizable, bodies possessing various degrees of solubility. They may be decomposed by nitrous acid giving off nitrogen and forming the corresponding hydroxy acid—*e. g.*,  $\text{CH}_2\text{OH.COOH}$ .

Amino-acids act both as bases and acids. The hydroxyl group in the carboxyl ( $\text{COOH}$ ) group makes it acid in reaction and it therefore reacts with metals to form amino salts of that metal; for example, zinc aminoacetate is



These salts also unite with other salts to form double salts. The basic properties are shown by the fact that amino-acids combine with acids to form double compounds like the following:



**Occurrence.**—Amino-acids occur normally in the urine. Any decomposition of meat by digestive ferments or bacterial action sets free amino-acids,

consequently they are found in the intestine and in the blood stream of the portal circulation. It will appear later that the extent of digestion and the place where it occurs can be studied by determining among other things the amino-acids.

#### SUMMARY OF CHAPTER XXXV.

The element nitrogen belongs in the same chemical group as arsenic, antimony, phosphorus, and bismuth. However, it has no physical properties in common with these elements.

Nitrogen is an odorless, colorless, inert gas. Four-fifths of the air consists of this inert element which serves to dilute the oxygen.

It is with considerable difficulty that nitrogen is made to combine with other elements. Even at high temperatures it resists union, though there are certain bacteria which grow on the roots of clover, beans and peas, capable of fixing atmospheric nitrogen. Electric sparks also are able to fix it. Under pressure at high temperatures in the presence of the proper activating agent, nitrogen combines with hydrogen to form ammonia,  $\text{NH}_3$ .

Ammonia gas dissolves in water to form ammonium hydroxide. This compound acts chemically like an alkali, forming salts with acids. Ammonia and ammonium hydroxide are caustic to the skin and mucous membranes.

Compressed ammonia, allowed to expand into large pipes, absorbs large amounts of heat. Advantage is taken of this property in refrigerating processes.

Ammonium salts appear as other crystalloids. They are very easily soluble in water and are volatile.

$\text{NH}_2$  is called the amino group. Nitrogen has a valence of three or five. The amino group has a valence of one, since two of the three bonds of the nitrogen are satisfied by hydrogen atoms.

An amino group may be substituted for one of the H atoms in the methyl groups of fatty acid. Such a substitution product is called an amino-acid. Amino-acetic acid has the formula  $\text{CH}_2\text{NH}_2\text{COOH}$ .

Amino-acids are crystalloids. They are decomposed by nitrous acid, setting free nitrogen, leaving the corresponding hydroxy acid as  $\text{CH}_2\text{OH}\text{COOH}$ . (hydroxy acetic acid or glycolic acid). Amino-acids act both as acids or bases, *i. e.*, they form salts and double salts with metals. They also combine directly with acids, *e. g.*, amino-acetic hydrochloride,  $\text{CH}_2\text{NH}_2\text{HCl}\text{COOH}$ .

Amino-acids occur in the products of digestion, in the blood and various tissues of the body, and in the urine.

Digestive processes of proteins may be followed to some extent by studying the liberation of amino-acids. Amino-acids are the chief constituents of proteins (egg white, casein of milk, gluten of wheat, etc.).

## CHAPTER XXXVI.

### OTHER NITROGEN COMPOUNDS.

**Cyanogen.**—If potassium carbonate and charcoal are heated together in a closed vessel in the presence of ammonia gas the following reaction takes place:



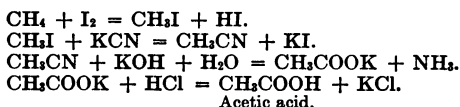
KCN is potassium cyanide, one of the most potent poisons. It is a white, crystalline solid, easily soluble in water. In water solutions it dissociates and the basic properties of the potassium are so much more prominent than the weakly acid reaction of the CN group that the solutions are alkaline in reaction.

In the presence of any mineral acid, like HCl, KCN is decomposed with the formation of hydrocyanic gas, HCN. This is a poisonous gas with a characteristic odor. Inhalation of the gas produces a choking sensation and dizziness—even small amounts may kill.

The group CN is called cyanogen. It is strangely like chlorine, bromine, iodine, and fluorine in its chemical behavior. The group combines with hydrogen as we have seen to form an acid; it also forms salts with metals and alcohol. It is monovalent. The

group does not exist alone—it does, however, exist as di-cyan (CN)<sub>2</sub>.

The cyanogen group offers a means for adding a carbon group to a radical. That is, by means of the CN group methane may be converted into an acid containing one C atom more than the alcohol.



**Sulphocyanates.**—If KCN is boiled with sulphur, potassium sulphocyanate, KCNS, is formed. This compound is of interest because it occurs in small amounts in the human saliva.

There are a great many other compounds of nitrogen with carbon, hydrogen, and oxygen which can not be discussed here.

**Anilin.**—Anilin has become a familiar word in recent years on account of the relation of this compound to dyestuffs. This substance was first obtained from the Indigo plant by distillation. Anilin is also found among the coal-tar products and may also be distilled from bones. The discovery of the graphic formula of this compound and its subsequent manufacture from benzene is one of the most brilliant examples of the use which chemistry has been to modern life. Its formula is C<sub>6</sub>H<sub>5</sub>NH<sub>2</sub>, which, on inspection, appears to be composed of an amino group replacing one of the hydrogen atoms in the benzene molecule. By chemical analysis this is found to be the case. In the preparation of anilin one cannot simply add an amino group to the benzol



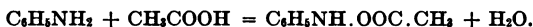
nucleus but must attain this end by some other route. First the substitution of the nitro radical is brought about by the action of nitric acid of benzene in the presence of sulphuric acid. This we will remember forms nitrobenzene,  $C_6H_5NO_2$ . The general method for getting rid of oxygen is by reduction, therefore, reducing the nitrobenzene by means of tin and hydrochloric acid, hydrogen combines with the oxygen, forming three parts of water and two other hydrogen atoms enter into the place vacated by the oxygen. Anilin is a colorless liquid which soon takes on a reddish color after exposure to air. It is used as a mordant in bacteriological work and is the substance from which a great many of our coal tar dyes are made. The synthetic production of these substances is so cheap in comparison with the cultivation and preparation of the indigo<sup>1</sup> plant that practically all of our dyes are made in the chemical laboratory.

Anilin is very sensitive to the action of reagents and for this reason offers many opportunities for the production of substitution products. It will be readily seen that its constitution is such that many substitution products are possible.

**Acetanilid.**—Anilin itself is a poison but its effect on the animal body can be lessened if other groups are attached to it. For example, if anilin is heated with acetic acid in an autoclave a crystalline compound is formed, which is used extensively in medicine. This

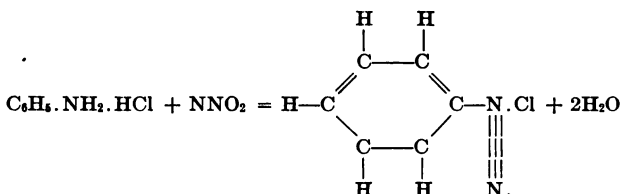
<sup>1</sup> Indigo dye is made synthetically but not from anilin. Anilin dyes are often used instead of indigo dyes.

compound is a product of the combination of anilin directly with acetic acid as follows:



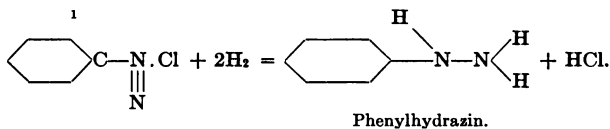
Administered in relatively small doses it is capable of reducing the temperature of patients with fever and is classed, therefore, among the febrifuges under the name of antifebrin.

**Diazonium Compounds.**—If anilin hydrochloride is treated with nitrous acid at low temperatures a *diazonium* compound is formed:



Observe that one of the nitrogens has a valence of three and the other a valence of five.

If the diazonium compound be reduced with tin and hydrochloric acid phenylhydrazin is formed.



Phenylhydrazin is a colorless liquid, with a high index of refraction. It combines with aldehydes and

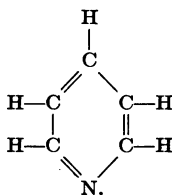
<sup>1</sup> The benzene nucleus is often represented by the hexagon. The six carbon and six hydrogen atoms with the three sets of double bonds are understood.

is therefore used in the partial identification of the sugars which contain this group. Sugars combine with phenylhydrazin to form characteristic needle crystals. The melting-point varies according to the groups of sugars entering into the compound. In practice in urine analysis, where it is desired to identify the reducing agent, the osazone is made. The osazone is a combination of two molecules of phenylhydrazin, with one molecule of sugar. Phenylhydrazin hydrochloride (colorless crystals) is generally used. An excess of this salt and sodium acetate are added to urine and boiled. On cooling, long, yellowish needles form if sugar is present. The needles are filtered off and their melting-point determined.

**Alkaloids.**—Such substances as strychnin, morphin, quinin, atropin, cocain, etc., are crystalline compounds which are obtained from different plants, and which constitute the active principle of these plants. For example, the physiological action of nux vomica is due in greater part to the presence of strychnin. These compounds have very complicated formulas. They are called alkaloids on account of the fact that they are like alkalies (bases) in forming salts with acids. The alkaloids themselves are very slightly soluble in water, but soluble in alcohol, while their salts are very slightly soluble in alcohol, but soluble in water. For this reason alkaloids are usually administered as salts rather than as free substances.

A great deal of the work has been done in establishing the graphic formulas of these substances with the result that we have definite knowledge of at least

six. These six alkaloids have been prepared in the laboratory, but on account of the expense attached to this method of preparation it has been found cheaper to obtain them from their natural sources rather than to manufacture them. The characteristic group of most alkaloids we may say is the pyridin ring:



The pyridin ring may be considered as a benzene ring with one of the CH groups replaced by N. This compound, pyridin, may be made synthetically by several processes. It is found in bone-oil.

**Vitamines.**—It has been found that certain nervous degenerations and disturbances may arise after the prolonged administration of certain diets. For example, beri-beri can be produced in birds by feeding them polished rice for a long period. Administration of small amounts of hulls or unpolished rice is sufficient to bring them back to normal; even minute amounts of extracts of the hulls or extracts of yeast seem to contain some substance or substances necessary for the normal processes of the body. These substances have not been obtained in pure form and consequently their chemical constitution is unknown. They have been termed, for want of a better name, "vitamines."

Experiments are being carried on in many places

in attempts to determine the constitution of these compounds. It will be noted that the first step in determining the constitution of the body is to obtain it in crystalline form so that chemical research is in the ultimate a search for crystals. So far the active principles classed under the heading vitamins have never been obtained in crystalline form, so that we know nothing of their constitution.

#### SUMMARY OF CHAPTER XXXVI.

Cyanogen,  $\text{CN}$ , is a group which acts very similarly to the halogens. It combines with hydrogen, with metals and alcohols. By means of it an organic radical may be oxidized into an organic acid containing one more carbon atom. Methane may be converted into acetic acid,  $\text{CH}_3\text{COOH}$ .

Anilin is amino benzene. It was first attained from the indigo plant but is now synthesized more cheaply than it can be extracted from the plant. From anilin a large number of dyes, drugs, perfumes, etc., are synthesized.

Anilin is a poison but when joined with an acetyl group its poisonous effects are greatly diminished. Thus acetyl anilid or acetanilid is used as an antipyretic.

A diazonium compound is formed by the action of nitrous acid on anilin hydrochloride. If this compound is reduced phenylhydrazine is formed. Phenylhydrazine is used to prepare osazones of sugars, whereby the latter may be partially identified. The melting-

point of the osazone gives a clue to the nature of the sugar.

Alkaloids are the active principles of plants. Most of them are crystalline substances very slightly soluble in water, but soluble in alcohol. They form salts with acids (hence their name alkaloid=like alkali) which are soluble in water and slightly soluble in alcohol. Alkaloids are usually administered as a salt. Six alkaloids have been synthesized but synthesis is more costly than extraction from plants. Much is known about several others.

"Vitamines" is a name given to a group of unknown substances which appear to be necessary for normal processes of the body. The feeding of polished rice produces a nerve degeneration known as beri-beri. Smaller amounts of the extract of yeasts, or alcoholic extract of rice hulls bring about recovery. No active crystalline compound has been isolated from the extracts. Therefore, these substances are not definitely known but merely suspected.

## CHAPTER XXXVII.

### PROTEINS.

THE most important nitrogen-containing substances constitute a large group of compounds similar in chemical structure and reactions, called proteins. The white of egg, casein (clot) of milk, and gluten of wheat are examples of proteins.

**General Characteristics.**—Proteins may apparently go into solution, but it has been found that they do not pass through membranes and collodion sacs. They are therefore in colloidal solution. Furthermore, some of them are completely coagulated by boiling with acids in the presence of salts. Both of these properties are characteristic of the colloids.

Coagulated proteins evidently undergo some sort of change for they cannot be redissolved into colloidal solution.

**Composition.**—The proteins are very complex in their structure. They are composed of the elements C, H, O, and N, generally sulphur and sometimes phosphorus.

Some proteins contain iron, and copper has been found in minute amounts in egg albumen. Iodine, zinc, and manganese have been said to be identified with the protein molecule.

The determination of the five principal elements occurring in proteins shows them to approximate one another very closely in empirical formulas at least. The principal proteins contain these elements in about the following percentages:

C about	52 per cent.
H about	6 per cent.
O about	25 per cent.
N about	15 per cent.
S about	1 to 2 per cent.

We have very little definite knowledge of the structure of the protein molecules. They are generally conceded to be composed of mono-amino-acids connected together in an unknown way. Some of the proteins contain aminodextrose (glucosamine). The protein molecule is very large but just how large has not been determined. Known methods for determining molecular weights are not applicable to proteins. Estimates of the molecular weight of egg albumen have been placed at 5000. The red blood coloring matter of the blood, oxyhemoglobin, is estimated to have a molecular weight of about 15,000. The simplest formula that can be calculated from analyses of oxyhemoglobin is  $C_{658}H_{1181}, N_{207}S_2FeO_{210}$ .

**Occurrence.**—Proteins are essential to life. The protoplasm (protein) of the cell is necessary for the continuance of life of the cell. While fats and carbohydrates can be replaced by one another and by proteins in the animal economy, the role of the proteins cannot be assumed by any other class of compounds.

Vegetable proteins are manufactured from simpler compounds by the plants, but the animal must depend



upon the vegetable kingdom for supply. The animal organism is capable of breaking down vegetable proteins or rebuilding them to suit the particular want, though it cannot build them from inorganic compounds.

**Varieties.**—On account of the lack of definite knowledge concerning the structure of proteins and on account of the large variety of them no classification free of exceptions has been made. Several classifications have been advocated by different associations of chemists but these serve only to confuse the student of elementary physiological chemistry. We may think of the proteins as a class as being made up of (1) natural or native proteids, and (2) changed or derived proteids (obtained from 1).

We may classify them also as: 1. Simple proteids. 2. Compound proteids. 3. Albuminoids.

*The Simple Proteids.*—White of egg, blood serum, and meat are examples of simple proteids. They are coagulated by heat in slightly acid solution if salts are present. They are precipitated (salted out) by saturation with ammonium sulphate; they form insoluble albuminates with  $\text{CuSO}_4$ ,  $\text{HgCl}_2$ ,  $\text{AgNO}_3$ , and  $\text{Pb}(\text{OOC}\cdot\text{CH}_3)_2$ . They give characteristic color reactions with certain reagents like nitric acid (xanthoproteic reaction, yellow) and alkaline copper sulphate (biuret, purple), etc.

The chief members of the simple proteids are albumins and globulins. Both of these proteids occur in blood serum.

*Globulins.*—If blood serum diluted with water is placed in a collodion sac which is allowed to float in

running water, the salts from the blood pass through the walls leaving the proteins behind. Within the sac a precipitate forms. This precipitate is globulin. After filtering this off, we find that globulin is not soluble in pure water but by the addition of some NaCl to the water, solution is accomplished. It is learned that salts are necessary for the solution of globulins. If NaCl or even  $MgSO_4$  is added to the globulin solution to the point of saturation the globulin is salted out. The same thing may be accomplished by adding an equal volume of ammonium sulphate (this makes, of course, only half saturation of this salt).

Globulin solutions answer the tests given above for simple proteids. Globulins are of interest on account of the fact that the valuable part of diphtheria antiserum is in the globulin fraction. Indeed, it is in this manner that diphtheria antitoxin is concentrated and much of the undesirable and invaluable part of the serum separated.

**Albumins.**—In the dialyzing experiment described under globulins, it was found that after the globulins had been separated there still remained in the serum solution a simple proteid.

This is the albumin fraction, which is soluble in water without the addition of any salts. Albumins are precipitated by full saturation with ammonium sulphate.

**Compound Proteids.**—Compound proteids are composed of a simple proteid joined to some non-proteid substance. They are classified according to the character of the non-proteid components, *e. g.*, *glyco-*

*proteids* yield on hydrolysis a simple proteid and a substance which reduces Fehling's solution. Mucin, the chief constituent of mucus (secretion of mucous membranes), belongs to the class of glycoproteids. Hemoglobin, the red coloring matter of the blood, yields a simple proteid plus hematin (an iron containing pigment).

The nucleins from the nuclei of cells is also a compound proteid.

*Albuminoids*.—Albuminoids are substances similar to the simple proteids. They occur in the horny layer of the skin, in cartilage, etc. Gelatin is formed by boiling the albuminoids found in cartilage (collagens) with dilute acid. Gelatin is a straw-colored jelly, when mixed with small amounts of water. It is soluble in water and not coagulated by boiling. It is used in making plates and tubes for bacterial growth.

*Derived Proteids*.—Protein substances form acid and alkali albuminates which are not coagulable by heat. They are changed chemically, especially the alkali albuminates. An alkali albuminate may be formed from an acid albuminate, but an acid albuminate can not be formed from an alkali albuminate.

**The Digestion of Proteins**.—Protein digestion is begun in the stomach. The contents of the stomach are kept acid by the secretion of hydrochloric acid from special glands in the stomach wall. The degree of acidity which these glands strive to maintain under normal conditions is equal to 0.2 per cent. HCl. The hydrochloric acid unites with the proteins to form acid albuminates, and the protein splitting ferment,

pepsin, now begins the disorganization of the protein molecule by breaking off amino-acids. The simplification of the protein molecule proceeds through the proteose stage until the protein is reduced to peptone plus amino-acids. Peptone is a simplified proteid (derived proteid) which is not coagulated by heat. Its molecule is smaller and it is easily soluble in water and diffusible. Peptone is precipitated from water solution by the addition of several volumes of absolute alcohol. Gastric digestion proceeds no further than the peptone stage and there is practically no absorption of protein digestion products from the stomach.

*Digestion of Protein in the Intestines.*—The acidified gastric contents pass into the small intestine and, meeting the bile and pancreatic juices, soon become alkaline in reaction. Alkaline albuminates are formed and the proteins are further simplified by trypsin, a protein splitting ferment secreted in the pancreatic juice. The results of the trypsin digestion of proteins are amino acids or compounds of two or more amino-acids.<sup>1</sup> The products of digestion are absorbed and transported to some unknown place by the blood and lymph to be rebuilt into proteins necessary for growth and repair. Proteins are also oxidized in the body, yielding 4 Calories per gram. It is not economical to furnish heat to the body in the form of proteins,

<sup>1</sup> The proteids are also broken up in the intestine to some extent by bacteria and the products are absorbed and detoxicated in the liver and excreted in the urine. Indican is one of these products. The bacterial decomposition of proteins are called *putrefaction* in contrast to *fermentation* which is the decomposition of sugars. So long as sugars are present in the intestines there is little putrefaction.

because of the higher cost of protein food and because of the increased metabolism when proteins are utilized.

Excess of protein is not stored in the body, but oxidized and excreted. Proteins are necessary for life and they cannot be replaced by carbohydrates or fats.

The amount of protein necessary for the maintenance of normal nutrition is at present a subject of discussion. There are those who believe that from 45 to 50 grams per day are all that are required for maintaining the body in health. However, if persons are left to select their own ration, they as a rule consume more than twice as much protein when it can be obtained. Experiments have been conducted, showing that the low protein diet is sufficient to nourish the normal individual when accompanied by sufficient amounts of carbohydrates and fats, but these experiments did not extend over a very long period and no question of resistance to disease was brought up. Too much protein can be digested and may result in hardening of the arteries, kidney lesions and probably functional disturbances, but a considerable excess of protein can be taken daily with safety. The trend of opinion seems to be on the side of the liberal allowance from 118 to 125 grams per day for the average man. Since we know very little of the mechanism of digestion and assimilation of these compounds and the relation of body proteins to those proteins found in our food, it is perhaps the wiser plan to allow the liberal amount of protein. While a person may remain in perfectly good health on a low protein diet, we are not sure that he will maintain his normal resistance against infecting agents.

## SUMMARY OF CHAPTER XXXVII.

Proteins are complex, nitrogen-containing compounds occurring as protoplasm, an essential part of the cell. Proteins found in plant cells have been synthesized from the elements. Animal proteins are derived from vegetable proteins or other animal proteins.

Proteins are compounds of C, H, O, N, S, Fe and sometimes P. Little is known of their structural formulas. They are made up of arrangements of amino acids. Amino glucose occurs in the molecule of some proteins. Most of our knowledge relates to the chemistry of the split products of proteins (amino acid compounds). Proteins may be classified according to their source: (1) natural or native proteids, and (2) changed or derived proteids (obtained from Class 1). They may be classified also as: (1) Simple proteids. (2) Compound proteids. (3) Albuminoids. To the first class belong white of egg, blood serum and meat. They are coagulated by heat (if salts are present), precipitated by acids, by saturation with salts and by certain chemicals like  $\text{AgNO}_3$ ,  $\text{HgCl}_2$ ,  $\text{CuSO}_4$ ,  $\text{Pb}(\text{OOC} \cdot \text{CH}_3)_2$ . All proteins are colloids, that is, they do not pass through semi-permeable membranes as salts do. A solution of proteins in a collodion sac placed in water will lose salts by dialysis into the surrounding water, but none of the proteins pass out. Salts are necessary to hold *globulins* (simple proteids) in solution. Globulins are found in blood serum. If diluted blood serum is dialyzed the globulin is precipitated. Globulin solutions are also precipitated by

half saturation with ammonium sulphate. Albumins (also simple proteids) are soluble in water without salts. Albumins are precipitated by complete saturation with ammonium sulphate. More than half of the proteins of blood serum consists of albumins.

Compound proteids are classified according to the non-protein components, *e. g.*, glycoproteids yield on hydrolysis a simple proteid plus a sugar. Mucin is a glycoproteid. The red coloring matter of the blood is a compound proteid, consisting of a simple proteid and hematin, an iron containing pigment.

Albuminoids are substances similar to simple proteids. They are colloids and contain nitrogen, but proteins can not be formed by the animal body from the products of their digestion. Cartilage contains albuminoids. Gelatin is a good example. Albuminoids are soluble in water and are not precipitated by heat.

Derived proteins are obtained from proteins by the action of chemical agents. Acids and alkalies unite with proteins to form, respectively, acid and alkali albuminates (derived proteids). They are soluble in water and are not coagulated by heat.

Protein digestion begins in the stomach. Acid albuminate is formed and broken down by pepsin to the peptone stage. Intestinal digestion takes place in alkaline solution. The alkaline albuminates are broken down to amino-acids by the trypsin of the pancreatic juice. The amino-acids are absorbed into the blood stream and, probably, in cells of the body new body proteins are formed from them. The excess

of proteins are changed into urea in the liver and excreted in the urine. Proteins are not stored in the body except in growth and repair. Proteins furnish heat (4 C. per gram) on oxidation in the body, but it is uneconomical to consume them for their heat-producing power. About 125 grams of protein per day is the average requirement. Proteins are determined quantitatively by estimating the amount of nitrogen by the Kjeldahl method and multiplying total N by 6.25.



## CHAPTER XXXVIII.

### THE BLOOD.

THE clot which forms when blood is drawn encloses all the cellular elements, leaving a clear straw-colored fluid known as serum. If the blood is drawn into oiled or paraffined tubes it will remain liquid for a long time. While liquid, the red and white cells and blood platelets may be centrifuged to the bottom, leaving straw-colored liquid, which appears to be serum. This is plasma. Plasma is serum plus the clotting substance or substances, for on standing or on coming into contact with any substance it can "wet" the plasma coagulates. The clot consists of fibrin and the remaining liquid is serum.

**Fibrin.**—Fibrin belongs to a class of coagulated proteids. It is the product of ferment action on fibrinogen. It is supposed that fibrinogen exists in the blood in solution. Fibrin-ferment is produced by the combination of calcium and an unknown body called pro-thrombin. Fibrin ferment + fibrinogen = fibrin. Fibrin forms a network of strands and includes the cellular elements of the blood in its meshes. The coagulation time is decreased with an increase in temperature. Fibrin may be obtained by whipping freshly drawn blood with a bundle of wires: the fibrin collects on the wires and leaves the cells of the blood suspended in the serum. If blood is drawn into sodium

oxalate or magnesium sulphate the calcium is precipitated and the blood will not clot. In practice blood is usually drawn into 2 per cent. sodium citrate solution in physiological salt solution. It does not clot, and the cells may be centrifuged down, the supernatant liquid poured off, and the cells washed with physiological salt solution.

**Blood Serum.**—As a rule blood serum is a clear, pale yellow to amber-colored liquid, although for some time after meals it may be whitish and turbid due to the fat globules in it. Serum is slightly more alkaline than plasma and has a specific gravity of about 1028 (1.028).

About 9 per cent. of serum consists of inorganic salts, most of which is NaCl. Potassium, magnesium, and calcium also occur. The alkalinity is due to sodium and potassium carbonate and sodium di-hydrogen-phosphate,  $\text{NaH}_2\text{PO}_4$ . About three-quarters of 1 per cent. (0.75 per cent.) of serum is proteid matter which consists of serum globulin and serum albumin. There is, perhaps, a little more albumin present than globulin.

The salts mentioned above are evidently necessary to keep the globulin in colloidal solution; if blood serum is placed in a collodion sac or animal bladder and this placed in a jar of pure water the crystalloids (salts) will pass through the wall of the sac<sup>1</sup> while the colloids and proteins will remain in the sac. But the globulin will be precipitated. The serum albumin remains in solution but can be precipitated by complete saturation with ammonium sulphate. The globulin fraction can

<sup>1</sup> This process is known as dialysis and has already been referred to under colloids.

also be precipitated by passing  $\text{CO}_2$  into diluted serum. This method is used in the purification and concentration of diphtheria antitoxin, since the antibody is contained in the globulin fraction.

**Blood Cells.**—The blood cells are mostly red blood corpuscles. For every 600 red cells there is normally one white corpuscle (leukocyte). The white cells are larger than the red cells and have a slightly lower specific gravity. The protoplasm of the white cells consists of conjugated proteid and globulin. Glycogen (animal starch) is also found in leukocytes. The white corpuscles have nuclei, while the red cells seem to be homogeneous (*i. e.*, constant composition throughout).

The red cells are very complex in chemical structure. Besides globulin, lecithin, etc., these cells contain the coloring matter of the blood, hemoglobin. Hemoglobin is a compound proteid containing iron and sulphur. It probably does not exist free in the red cell but in combination with other proteins. It is of interest here because it plays the chief role in internal respiration. Oxygen forms a loose combination with this proteid called oxyhemoglobin, is transported to distant parts of the body, and given up where needed. Carbon monoxide,  $\text{CO}$ , forms a stable compound with hemoglobin, which cannot be broken up without disintegrating the hemoglobin.

**Osmosis.**—The red corpuscles may be suspended in serum or in 0.9 per cent.  $\text{NaCl}$  solution without undergoing appreciable change. If, however, water is added to the mixture the cells may be seen under the microscope to swell and finally burst. The suspension of

red cells loses its turbidity, becomes clear and pink in color, due to the solution of the red coloring matter. The swelling of the red cell is the result of the entrance of water into it—osmosis. The same phenomenon can be observed by placing vegetable cells in different concentrations of salt solutions. The experiment can be carried out on a large scale by filling a bladder with a solution of salt and placing it in a vessel of water. If the bladder is tightly closed it will finally burst on account of the inflowing of water. If a tube had been placed in the neck of the bladder so that water could rise in it, the pressure of the incoming water can be measured. This is called osmotic pressure. Osmotic pressure may reach an incredible figure: over 30 atmospheres (400 pounds to the square inch) has been observed.

Many phenomena of cell life are explained on the basis of osmosis.

It is interesting to note that molecular weight may be determined by observing the osmotic pressures under certain conditions.

**Tonicity.**—Solutions having the same molecular concentration are said to possess the same tonicity or are isotonic. Thus, 0.9 per cent. NaCl is isotonic with blood serum, *i. e.*, blood cells placed in 0.9 per cent. NaCl will not burst. If water is added and the percentage of NaCl is therefore reduced the tonicity is lowered and the solution is said to be hypotonic (less than tonic). If too much salt is present the solution is hypertonic (greater than tonic). A hypertonic solution causes the red cells to shrink (crenation). We thus see the

reason for preparing physiological salt solution exactly 0.9 per cent. NaCl. It has already been stated that physiological salt solution should not be called *normal* salt solution because a *normal* (chemical) NaCl solution is a 4 per cent. solution. Isotonic (for blood cells) salt solution is about seventh molecular, M/7.

**Hemolysis.**—The dissolution of red cells by water is known as lysis, hemolysis or laking. There are specific poisons (snake venoms) and natural agents in blood from other species which are hemolytic. Hemolysins (agents causing hemolysis) can be called forth in a serum by the intravenous injection of red corpuscles of other animals. For example, sheep corpuscles injected intravenously into rabbits call forth a hemolysin capable of destroying sheep corpuscles. This principle is used in the Wassermann reaction.

Natural hemolysins potent for human red cells sometimes occur in members of the same species. It is therefore necessary to test the serum of the donor against the red cells of the patient before transfusions.

**General Characteristics of the Blood.—The Amount.**—About one-fourteenth of the body weight of man is blood. In smaller animals the amount is proportionally greater. The average 150 pound man has about 11 pounds of blood or about five quarts.

**Specific Gravity.**—The specific gravity of blood may be determined by placing drops of it in a mixture of chloroform and benzol. If the drop rises add benzol, if it sinks add chloroform until it comes to rest about halfway. The specific gravity of the mixture can then be determined with a hydrometer—this figure would

be the specific gravity of the blood. The average specific gravity is 1056 (1.056).

**Functions of the Blood.**—The blood is the great equalizer of the body. It equalizes heat and moisture, it brings food and oxygen to the cells and takes away  $\text{CO}_2$  and refuse material. It brings the protective agencies to a point of injury and invasion and maintains constant guard over all portions of the body. Its composition is subject to extreme variation.

Besides gases, the blood holds liquids in solutions and solids in solution or suspensions, and it carries poisonous products to the liver for detoxication, and then again transports them to the kidneys for elimination. Urea, uric acid, ammonia, etc., found in the urine are brought to the kidneys by the blood.

The blood receives internal secretions from the ductless glands like the adrenals, the pancreas, the pituitary body, and the thyroid gland.

Dextrose is found in the blood under normal conditions in about one to one and a half parts per thousand. Whether or not it is held in loose combination with some substance is not known; but when the concentration reaches 3 parts per 1000 sugar appears in the urine. The sugar content increases during digestion.

Various ferments are found in blood. Among them are found oxidizing and fat splitting ferments. The presence of the oxidizing ferment offers a very delicate method for the detection of blood. Blood as dilute as one part in eighty million will give a positive peroxidase test.

## SUMMARY OF CHAPTER XXXVIII.

Blood drawn into paraffined tubes will not clot immediately. The cellular elements may be removed, leaving a clear, yellow, alkaline fluid called plasma. Plasma clots on coming into contact with any foreign body it can "wet." Blood allowed to clot carries down the cellular elements in the meshes of the clot leaving a clear, yellow, alkaline, sticky liquid called serum. The clotting substance is fibrin. It may be obtained from the fresh blood by whipping with wires. The fibrin sticks to the wires.

Fibrin exists in the blood as fibrinogen. A hypothetical substance called prothrombin unites with calcium to form fibrin ferment which in turn acts on the fibrinogen to produce fibrin. Calcium is essential to the clotting of blood consequently anything which removes the calcium as sulphates or oxalates prevents the clotting of blood.

Blood serum consists of a colloidal solution of globulin and albumin in an alkaline solution of salts. The salts consist mostly of  $\text{NaCl}$ ,  $\text{MgCl}_2$ ;  $\text{KCl}$  and  $\text{CaCl}_2$  are also present. The alkalinity of the serum is due to the presence of alkali carbonates and dihydrogen phosphates. Globulin may be precipitated by removing the salts by dialysis or half saturation with ammonium sulphate. Albumin may be precipitated by complete saturation with ammonium sulphate. The ratio of globulin to albumin is about 1 to 1.5. Globulin is also precipitated by diluting the serum and passing  $\text{CO}_2$

into the mixture. Diphtheria antitoxin is thus concentrated.

The red cells are complex in their composition. Besides globulin, lecithin, etc., they contain hemoglobin, a compound proteid. Hemoglobin consists of simple proteids plus hematin, the red coloring matter of the blood. Hemoglobin unites with oxygen in the lungs to form oxyhemoglobin and gives up the oxygen to the various cells of the body which need it.

If a capsule of semipermeable membrane containing salt solution is placed in water, the water passes into the salt solution through the wall. This is called osmosis. The pressure exerted in the process is called osmotic pressure. Solutions containing equal molecular concentrations of salts and exerting the same osmotic pressure are said to be isotonic. More dilute solutions would be termed hypotonic and more concentrated solutions hypertonic. A solution of 0.9 per cent. NaCl is isotonic with the red cells and therefore with blood serum. More dilute solutions burst (hemolyze or lake) the cell and the coloring matter is dissolved out. Poisons and natural substances found in blood, also substances developed after the repeated injection of washed red cells, which hemolyze red cells, are called hemolysins. Monkey serum possesses a strong natural hemolysin for sheep corpuscles. The serum of some persons will hemolyze the red cells of others. Therefore, before a blood transfusion, the two bloods (of the donor and of the patient) should be tested against one another for hemolysins.



About one-fourth of the body weight is blood. The specific gravity of blood is about 1056.

Blood carries oxygen to the body cells and brings away  $\text{CO}_2$ . It also furnishes food and water to the cells and removes their waste products. Blood also equalizes the temperature of the body and maintains a guard over the various parts. When local injuries or infections occur the blood transports the white cells and neutralizing bodies to the point.

Dextrose is found in normal blood (1 to 1.5 parts per 1000).

Ferments of various kinds are also found in normal blood.

## CHAPTER XXXIX.

### MILK.

MILK is an albuminous fluid in which fat is emulsified and the protein (casein) is held in colloidal solution. If one adds rennet (a ferment taken from the walls of pig's stomach) to milk, it clots. The casein settles out in masses, leaving a slightly yellow, clear fluid above. Acids added to milk produce the same effect but in a different manner. It is thought that the casein is held in colloidal solution by calcium, and any substance which takes away the calcium allows the casein to clot. For example, hydrochloric acid added to milk unites with the calcium, taking it away from the casein. Calcium chloride is formed and casein is precipitated. The clot may be washed and shaken with lime water to reform a solution of casein. The souring of milk is the result of the growth of acid-producing microorganisms in it. The lactic acid formed by the fermentation of milk sugar (lactose) acts in the same manner as the hydrochloric acid referred to above.

**Casein.**—Casein is a compound protein, very complex in structure. In the milk it is said to exist as caseinogen, which is still more complex. Caseinogen is insoluble in water and so is calcium phosphate, but

together they form a colloidal solution which gives to milk its white appearance.

Besides its use as a food, casein is employed extensively in making artificial ivory. Milk is the only source of casein.

**Other Proteins in Milk.**—There are two proteins found in milk which are very closely related to those found in blood, viz.: lactalbumin and lactoglobulin. Milk secreted the first few days after delivery (colostrum) is relatively richer in these two proteins than normal. Human milk is proportionally richer in lactalbumin and lactoglobulin than cow's milk.

**Fats.**—Milk fats are emulsified in milk but on standing the globules gradually rise on account of the difference in specific gravity. Forcing the milk through fine holes in a steel plate homogenizes it so that the cream will not rise.

When milk is churned the violent agitation destroys the emulsion and the fat globules coalesce to form butter. The fat globules of human milk are smaller than those found in cow's milk. There is also some difference in the chemical composition (see table).

**Milk Sugar.**—Milk sugar, lactose, is found only in milk. It is composed of a molecule of dextrose and a molecule of galactose. Human milk contains about 7 per cent., and cow's milk about 5 per cent., lactose.

**Salts.**—The salts in milk are very similar to those found in the blood. The kinds and relative amount of salts determine the reaction of fresh milk. Human milk is more alkaline than cow's milk.

## HUMAN AND COW'S MILK.

Milk from every source is similar in its constituents, but it may vary greatly in different species. A comparison of human and cow's milk is given in the following table:<sup>1</sup>

Reaction.	Woman's milk direct from the breast. Amphoteric (more al- kaline than acid).	Cow's milk, freshly milked. Amphoteric (more acid than alkaline).
Water . . . . .	87 to 88 per cent.	86 to 87 per cent.
Mineral matter . . .	0.20 per cent.	0.70 per cent.
Totals solids . . . .	13 to 12 per cent.	14 to 13 per cent.
Fats . . . . .	4.00 per cent. (rela- tively poor in vola- tile glycerides).	3.50-4.00 per ct. (rela- tively rich in volatile glycerides).
Milk sugar . . . . .	7.00 per cent.	4.75 per cent.
Proteids . . . . .	1.50 per cent.	3.50 per cent.
Caseinogen . . . . .	$\frac{1}{2}$ to $\frac{1}{3}$ of total proteids.	2.66 per cent.
Whey proteids . . . .	$\frac{1}{3}$ to $\frac{1}{2}$ of total proteids.	0.84 per cent.
Coagulable proteids .	Small proportionally.	Large proportionally.
Coagulation of pro- teids by acids and salts.	With greater diffi- culty. Curds small and flocculent.	With less difficulty. Curds large and tenacious.
Coagulation of pro- teids by rennet.	Does not coagulate readily.	Coagulates readily.
Action of gastric juice	Proteids precipitated but easily dissolved in excess of the gas- tric juice.	Proteids precipitated but dissolved less readily.

**Milk as a Food.**—It has been said that milk is the ideal food, containing as it does all three classes of foodstuffs, carbohydrate, fat, and protein. It will be noted that while milk is an ideal food for the young and growing child, it is not a balanced ration for older children or adults. It has far too much protein and contains too much fat in proportion to the amount of sugar present. Furthermore, it is not sufficiently

<sup>1</sup> Quoted from Rotch: Pediatrics.

concentrated. In order to obtain a normal amount of heat for the metabolism of a normal man during one day, it can be calculated that over four quarts of milk would be necessary. Milk offers a very convenient food for invalids and for persons requiring such protein diets, and skimmed milk is one of the cheapest sources of protein for the normal dietary. Milk is the most difficult foodstuff to produce and transport in a hygienic manner. All sorts of microorganisms flourish in it, and several diseases have been known to have been transmitted in epidemic form by milk.

#### SUMMARY OF CHAPTER XXXIX.

The casein of milk is held in colloidal solution by calcium phosphate. In milk solution it exists as caseinogen, which, alone, is insoluble in water. Calcium phosphate is also insoluble, but caseinogen and calcium phosphate together, form a compound which goes into colloidal solution. Obviously, anything which removes the calcium from this combination precipitates the caseinogen as casein. Acids added to milk unite with the calcium and the casein clots. The souring of milk is the result of the growth of microorganisms which produce lactic acid from the milk sugar. Lactic acid forms calcium lactate and the casein is precipitated. Rennet, a ferment found in the wall of the pig's stomach, has as its particular function the clotting of milk.

Casein is a complex compound protein. 2.66 per cent. of cow's milk or 0.75 per cent. of human milk is casein.

The clear liquid (whey) resulting from filtering clotted milk is alkaline in reaction and contains practically the same inorganic salts as blood serum, but in different proportions. Whey also contains lactalbumin and lactoglobulin. There are relatively larger amounts of lactalbumin and lactoglobulin in human milk than in cow's milk. Colostrum contains far greater proportion of these two proteids.

The fats of milk are emulsified in milk but on standing rise to the top (cream). About 4 per cent. of normal milk is fat, but cow's milk often contains less. The globules in cow's milk are larger and there are chemical differences between the fat of cow's milk and that of human milk.

Milk sugar, lactose, is composed of a molecule of dextrose plus a molecule of galactose. Lactose is found only in milk. Human milk contains about 7 per cent., and cow's milk about 5 per cent., lactose.

Milk contains all the necessary food materials, namely, protein, fat, sugar, but not in the proportions required by adults. Milk is the ideal food for babies, but contains too much protein to serve as an exclusive diet for adults. It is, however, a very economical source of protein for the dietary. Milk is the most difficult foodstuff to produce in a hygienic manner.

## CHAPTER XL.

### THE URINE.

URINE is a solution of various salts and certain organic nitrogenous bodies in water. On its way through the urinary passages it may wash down particles consisting of debris and cells. In certain pathological conditions the amount of matter not in solution is increased. On cooling, certain crystals may be deposited, and if the reaction changes a visible precipitate may form.

**Amount.**—The amount of urine excreted varies very considerably. More is excreted during activity and just after meals. Coffee and tea often increase the flow temporarily. Obviously the amount of liquids and the kind and amount of food taken affect the volume of the urine. The temperature of the air and humidity have very marked effects. When it is hot, and both sensible and insensible perspiration are increased, the urine is more concentrated than on colder days. The average urine output varies from 800 to 1200 c.c. for the average man in twenty-four hours. In certain diseases the urine is increased, while in other disorders the reverse is true.

**Specific Gravity.**—Since the urine contains salts and organic bodies in solution, the specific gravity would be

expected to be greater than 1. The variation in volume under normal conditions would be expected to cause some variation in the specific gravity. The specific gravity ordinarily varies between 1015 and 1025 (or 1.015 to 1.025). In clinical work it is customary to express the specific gravity of water as 1000, and of urine in terms of this number. After excessive perspiration it may go as high as 1040, and after a meal it may be as low as 1010. In very rare instances does it go below 1000—probably then only after excessive amounts of alcohol have been taken. The specific gravity is estimated by means of the urinometer. (See page 250.)

**Color.**—Normal urine when excreted is always clear, but it has a distinct color, varying from a pale lemon yellow to deep orange. The normal yellow color most often found is due to the urine pigment: urochrome, a substance whose constitution is not yet determined. The ingestion of drugs and vegetable coloring matters may affect the color of the urine. The depth of color will, of course, depend upon concentration and vary inversely as the volume, other things being equal. In pathological conditions bile or blood, etc., may lend color to the urine.

**Odor.**—Very recently it has been claimed that the faint aromatic odor of fresh urine is due to the presence of a particular substance termed uronoid. Its constitution has not been determined. After the ingestion of asparagus a characteristic odor is observed due, it is claimed, to the presence of a methyl mercaptan (a sulphur-nitrogen-carbon containing organic



compound). Other vegetables and drugs may impart odor to the urine. Urine decomposes very quickly and its odor becomes very unpleasant. Ammonia is given off in sufficient quantities to be detected by its odor.

**Reaction.**—Normal human urine is usually slightly acid to litmus, due to the presence of sodium dihydrogen phosphate,  $\text{NaH}_2\text{PO}_4$ . Immediately after meals when hydrochloric acid is being secreted in the stomach the urine may be slightly alkaline because the HCl secreting glands extract the acid radicals from the blood for the time being. Organic acids, like acetic (in vinegar) and citric (in lemon juice), are oxidized in the body to carbonates. The carbonates are excreted in the urine and consequently increase the alkalinity.

On standing, fermentation results in the formation of ammonia and the urine becomes alkaline. (See above.) Alkaline urine becomes cloudy on account of the precipitation of phosphates. A few drops of acetic acid will clear it (*i. e.*, dissolve the phosphate). Alkaline urine may deposit small crystals of ammonium urate and ammonium magnesium phosphates.

**Salts of the Urine.**—The salt occurring in the urine in largest amounts is sodium chloride. The actual amount eliminated from day to day depends upon the amount ingested upon the amount excreted in the perspiration. The average is about 12 grams per day. Chlorides of potassium, ammonium, and magnesium also occur in the urine. The phosphates of urine are also variable. Phosphates of the alkali

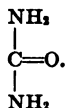
metals (Na and K), of the alkaline earths (Ca and Mg), and of ammonium are found. The three types of phosphates may be represented by the sodium salt: Normal phosphate,  $\text{Na}_3\text{PO}_4$ ; monohydrogen,  $\text{Na}_2\text{HPO}_4$ ; and the dihydrogen  $\text{NaH}_2\text{PO}_4$ . On account of the low solubility of phosphates they are often found deposited as microscopic crystals in cold urine. The sulphur of the urine occur as inorganic sulphates and also in organic combinations (ethereal sulphate and neutral sulphur compounds). Small amounts of carbonates and nitrates, traces of fluorides and silicates, and some iron are also found in urine in addition to the salts already enumerated.

**Organic Constituents.**—The most important organic constituent of the urine is urea. About 90 per cent. of the total nitrogen of urine is urea. An average of 30 to 40 grams of urea are excreted by man in twenty-four hours.

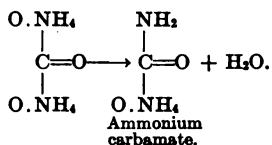
If nitric acid is added to urine and the mixture evaporated to one-third volume and set aside to cool crystals resembling six-sided tiles will form. These crystals will be colored, but if they are filtered off, dissolved in water, filtered through powdered animal charcoal and concentrated they will crystallize without color.

These crystals are urea nitrate. If  $\text{BaCO}_3$  is added to a solution of urea nitrate,  $\text{CO}_2$  is given off,  $\text{Ba}(\text{NO}_3)_2$  is formed and urea is set free. Urea is soluble in water but insoluble in chloroform or ether. It crystallizes in colorless, long, six-sided prisms.

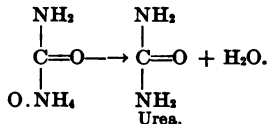
Urea has the composition



It can be made synthetically by heating ammonium carbonate. The first step is the driving off of 1 molecule of water forming ammonium carbamate:



Continued heating changes ammonium carbamate in the same manner.



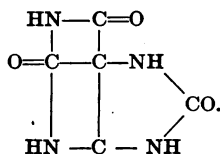
Urea is decomposed by  $\text{HNO}_2$  (nitrous acid) into gaseous nitrogen and  $\text{CO}_2$ . The  $\text{CO}_2$  can be absorbed by  $\text{NaOH}$  and the amount of N gas measured. The amount of urea can then be calculated. (See chapter on Uranalysis.)

It has been stated (see proteins) that proteins are broken down during digestion into amino-acids. The amino-acids are absorbed and at some unknown place are reassembled to form proteins when they are necessary for growth or repair. Excess of proteins are not stored in the body as fat and sugars are. The excess of amino-acids are oxidized in the liver to form urea and

excreted in the urine. Therefore the amount of urea in the urine varies with the amount of proteins ingested. Urea is also excreted in the sweat.

*Uric Acid.*—Although the actual amount of uric acid excreted in twenty-four hours is comparatively small (about 0.5 gram) it is considered an important constituent of the urine. It is supposed to come from the breaking down of the nuclei of body cells and from the nuclei of cells in the meat ingested.

The crystals of uric acid may assume several forms: like wedges, prisms, plates, or dumb bells. Free uric acid is insoluble in water but its lithium and its sodium salts are easily soluble. Ammonium urate is very slightly soluble. Crystals of uric acid or ammonium urate are colorless when pure, but they occlude coloring matters so that often the brownish-red precipitate seen in urines, after standing, consists of uric crystals or its salts. Uric acid in excess reduces Fehling's solution slightly, and may, therefore, when present in large amounts, interfere with sugar tests in the urine. The formula for uric acid is:



It is a tri-oxy-purine. On oxidation the three C atoms connected with one another are changed to  $\text{CO}_2$  and two molecules of urea are given off.

*Indican.*—Indoxyl potassium sulphate or indican is one of the most important of the ethereal sulphuric

acids found in the urine. It is oxidized by chlorine (bleaching powder) to indican blue or indican red. (See test for Indican in Appendix.) Indican is a product of putrefaction and when found in the urine is evidence that putrefaction is taking place in the intestinal contents.

Creatinin seems to be a product of cell activity, and on a creatin-free diet the amount excreted seems to be fixed for each individual. Creatin occurs in muscle: creatinin is reduced creatin. For formula and properties, refer to any text-book on physiological chemistry.

**Other Organic Constituents of the Urine.**—A great many other organic compounds of more or less complexity may occur in variable quantities. Their presence in small quantities may be normal, but the extent to which they may be present to constitute abnormality varies under different conditions.

## SUMMARY OF CHAPTER XL.

The amount of urine excreted varies markedly. The temperature and humidity, the amount of water drunk, the kind and quantity of food ingested, and the blood-pressure affect the secretion of urine. The average varies from 800 to 1200 c.c. per day.

The specific gravity varies with the volume and with diseased conditions. It generally varies from 1015 to 1025, though it may exceed these limits under certain conditions and still be normal. Under normal conditions, urine of high specific gravity makes us

suspicious of diabetes and tests for sugar in urine are accordingly made.

The color is also subject to considerable variation from pale yellow to deep orange. The normal color is due to urochrome. Bile salts are found in abnormal conditions. Substances taken into the stomach may be eliminated in the urine and change its color.

The odor of the urine is said to be due to a substance of unknown composition called uronoid. Substances in the food may impart a characteristic odor to the urine, *e. g.*, asparagus.

The reaction of fresh, normal, human urine is slightly acid, although after the ingestion of large amounts of vegetables or organic acids it may be alkaline. Decomposition of various nitrogenous bodies in the urine by ferment action soon renders the urine alkaline. This alkalinity results from the ammonia liberated.

The salts of the urine consist mainly of NaCl. About 12 grams of NaCl per day are excreted. Chlorides of K, (NH<sub>4</sub>), Mg, and Ca are also found. Phosphates of alkaline metals and alkaline earths occur. The three types of phosphates found may be represented as follows: (1) M<sub>3</sub>PO<sub>4</sub>; (2) M<sub>2</sub>HPO<sub>4</sub>; and (3) MH<sub>2</sub>PO<sub>4</sub>, in which M represents any monovalent metal like Na or K. Sulphates (inorganic and organic) and small amounts of carbonates and nitrates are found.

The most important organic constituent is urea. 90 per cent. of the total N is urea N. 30 to 40 grams of urea are excreted by an average man in one day. Urea can be separated from urine by adding HNO<sub>3</sub>. The urea nitrate can be recrystallized and finally decomposed by BaCO<sub>3</sub>. Urea has the composition

$\text{CO}(\text{NH}_2)_2$ . It can be made synthetically by heating ammonium carbonate. Ammonium carbamate  $\text{CO}.\text{NH}_2.\text{O}.\text{NH}_4$  is first formed and later changes to urea on losing 1 molecule  $\text{H}_2\text{O}$ . Urea is formed in the liver by the oxidation of the groups of amino-acids which have not been utilized to rebuild proteins.

Urea is decomposed by  $\text{HNO}_2$  forming gaseous nitrogen and  $\text{CO}_2$ .

Uric acid is supposed to result from the breaking down of the nuclei of ingested cell nuclei. On standing, urine deposits crystals of uric acid, especially after the urine becomes alkaline. Uric acid forms salts with Na, K, and  $(\text{NH}_4)$  groups. Sodium and potassium urates are much more soluble in water than ammonium urate. Uric acid in excess reduces Fehling's solution and may thus be a source of error in examining a sample of urine for sugar. Uric acid is a tri-oxy-purine (closely related to the alkaloid caffeine). Urea is formed by the oxidation of uric acid.

Indican is indoxyl potassium sulphate. When this body is found in urine it indicates intestinal putrefaction. It may be detected in urine containing it by treating with bleaching powder. A blue or red color indicates the presence of indican.

Creatinin seems to be a product of cell activity, and on a creatin-free diet the urine of individuals contains a fairly constant amount of creatinin peculiar to that person. Creatin occurs in muscle: creatinin is reduced creatin.

Other organic compounds may be found at times in the urine, but are of interest more to the students of physiological and pathological chemistry.

## CHAPTER XLI.

### URANALYSIS.

THE normal appearance of a sample of urine is no guarantee that it is normal. It is very important that specimens should be tested, even though they appear to be perfectly normal.

**Collection of the Urine.**—For chemical analysis of urine it is extremely important that a well-mixed twenty-four-hour specimen be collected, for the reason that the character of the various voidings are subject to wide variations.

Sometimes it is desired that the night urine be kept separate. In this case the voidings between 9 P.M. and 6 A.M. are poured together. A twenty-four-hour specimen includes voidings beginning with the first after 6 A.M. and including the 6 A.M. voiding on the next day. It is very essential that all voidings should be saved and the total accurately measured.

**Preservation of Urine.**—All specimens should be kept in the ice-box until immediately before examination. Urine is easily fermented—the reaction changes, urea is decomposed, the ammonia content increases, carbohydrates may be fermented, and the microscopic elements decompose if the specimen is not kept cold.

Various preservatives have been recommended for



special kinds of work. The analyst should know the preservation used. For chemical work chloroform is perhaps the best. Add 5 c.c. for each liter of urine. The urine should be well shaken after each addition and the bottle kept tightly corked. The chloroform settles to the bottom and adds nothing to the volume. Immediately before analysis the urine may be poured off the chloroform and a sample exposed to incubator temperature for a very few minutes, or air bubbled through in order to remove traces of chloroform. Any larger trace of chloroform may yield a suggestive reducing reaction with Fehling's solution.<sup>1</sup> If formalin is used one is more liable to obtain suggestive sugar reactions with Fehling's solution, yet for microscopic examination formalin seems the best preservative. Thymol is used sometimes, but a test similar to bile may be obtained in the specimen.

**Amount of Urine.**—The amount of urine should be carefully measured in a cylinder registering volume in cubic centimeters.

**Specific Gravity.**—For clinical examination the specific gravity is approximated by using the urinometer. The instrument is standardized to read 1000 in distilled water at 15° C.

The glass cylinder should be clean and dry. It is filled about four-fifths full of urine by gently pouring in order to prevent foaming. Foam should be removed with a strip of filter paper.

<sup>1</sup> Formalin does not reduce Benedict's solution but chloroform does. For each liter of urine 2.5 c.c. formalin is sufficient.

The temperature is important. A small thermometer should be placed in the urine and stirred a few times. If the temperature is not 15° C. the jar should be placed in warm or cool water, as the case requires, and gently stirred until the thermometer registers 15°. If the proper temperature cannot be attained in a hurried examination, a correction by adding for every 3° above 15° is made. For example, a urine reading 1022 at 21° should be corrected to read 1024.

The bobbin should be clean and dry. It is inserted carefully and gently tapped and read when it comes to rest. It is important that the bobbin does not touch the side or bottom of the cylinder. The graduation on level with the surface of the urine (the bottom of the meniscus) is read and recorded.

In a child's or catheterized specimen there may not be sufficient urine to fill the cylinder. Dilute with equal parts of water and test as above. Multiply the last two figures by 2 and add to 1000. If reading is 1010, then  $10 \times 2 = 20$ ;  $20 + 1000 = 1020$  (correct reading).

**Total Solids.**—When requested, estimate total solids by multiplying the last two figures of the specific gravity by Häser's empirical coefficient 2.33. This will give grams per liter. To find percentage divide by 10.

**Color.**—The color of a given depth of urine viewed against a white background is recorded in well-recognized terms: straw, light yellow, yellow, amber, orange, brown, brownish red, etc. If the foam is yellow, bile is present. The foam is otherwise white.

The character of any precipitate is described as to amount (small, moderate, abundant), color, and consistency (dense or flocculent). The general appearance of the urine (clear, cloudy) is observed.

**Odor.**—A small amount of urine is brought to boiling in a beaker covered with a watch-glass. The glass is removed and odor recorded as fruity, ammoniacal, sulphidic, etc.

**Reaction.**—A strip of blue and of red litmus paper are immersed to half their length in the urine and immediately withdrawn. The color change is observed. If the blue paper changes quickly to red, the urine is strongly acid; a slight change is interpreted as faintly acid. The strips are laid on white, clean, dry filter paper. If the red paper changes blue, the urine is recorded as being alkaline. If the strip of litmus is allowed to dry and becomes red again, the alkalinity is recorded as being due to ammonia, otherwise it is due to a fixed alkali. Normal urine is always faintly acid, varying with the diet.

**Total Acidity.**—When required, the total acidity may be determined by titrating with  $\frac{n}{10}$  NaOH, according to the method of Folin.

25 c.c. of urine are delivered from a pipette into a 200 c.c. Erlenmeyer flask. 100 c.c. distilled water, 20 grams potassium oxalate, and 2 drops of a 0.5 per cent. alcoholic solution of phenolphthalein are added. The flask is shaken well for one minute and  $\frac{n}{10}$  NaOH added from a burette until a distinct reddish color is produced. The flask is shaken after each addition. The amount of NaOH necessary to produce color when

multiplied by 4 gives the number of c.c.  $\frac{N}{10}$  alkali required per 100 c.c. urine.

**Albumin.**—Albumin is coagulated and precipitated by heat. Pour into a clean chemical test-tube 2 c.c. of a saturated solution of NaCl in water. Pour in filtered, clear, slightly acid urine until three-quarters full. The top zone of  $\frac{3}{4}$  to 1 inch is heated to boiling in a small flame (alcohol lamp preferred). If a whitish cloud is observed in the heated zone, when viewed against a black background, add one drop of 25 per cent. acetic acid and boil again. Repeat the addition of acetic acid and the boiling three times. If the precipitate is due to phosphates this treatment will cause it to disappear. If the cloudiness is due to albumin it will persist and perhaps increase. The whitish deposit sometimes formed on the glass by the flame, especially when gas is used, should not be confused with an albumin test. If much albumin is present a heavy flocculent precipitate will be formed.

**Heller's Nitric Acid Test.**—Place 10 c.c. of clear, concentrated nitric acid in a two-ounce conical stand glass. (If the nitric acid is yellow, boil in a beaker until clear.) With a pipette carefully overlay about 25 c.c. filtered urine over the nitric acid. After three minutes observe the line of contact against a dark background. A precipitate at the juncture of the two liquids indicates albumin. The red or reddish-violet ring often observed should not be mistaken for a precipitate.

**Reducing Sugars.—Fehling's Method.**—*Solution A.*—Dissolve 34.65 grams powdered  $\text{CuSO}_4$  in 300 c.c. warm, distilled water in a 500 c.c. volumetric flask.

Allow to cool and make up to mark (500 c.c.) with cold, distilled water.

*Solution B.*—Weigh out roughly on filter paper 125 grams potassium hydroxide. (KOH is caustic; handle with forceps. Break sticks in a large mortar.) Place in a 500 c.c. beaker and add 300 c.c. distilled water. When all is dissolved pour into a 500 c.c. volumetric flask and rinse three times with small portions of distilled water, pouring each portion into the flask. Pulverize about 180 grams sodium-potassium tartrate (Rochelle salt) in a mortar. Weigh out exactly on glazed paper 173 grams of powdered Rochelle salt and pour carefully into flask containing the KOH solution. Shake until all is dissolved; allow to cool and make up to 500 c.c. mark with distilled water; preserve in a rubber-stoppered bottle.

For tests, mix equal parts of solutions *A* and *B* immediately before using. To 1 c.c. of the mixture in a test-tube add 4 c.c. distilled water and boil. (If a precipitate is formed the solutions are worthless.) To the warm solution add three or four drops of the urine to be tested and boil. Repeat this several times. A yellow or brownish precipitate indicates the presence of reducing sugars. Phosphates and uric acid may form precipitates which may be confused with the positive reaction. Fehling's solution should not be used after it has stood over long periods.

**Benedict's Method.**—Dissolve 173 grams sodium citrate and 100 grams anhydrous sodium carbonate in 600 c.c. distilled water. Pour through a folded filter into a glass graduate and make up to 850 c.c. with

water. Dissolve 17.3 grams powdered copper sulphate in 100 c.c. water and make up to 150 c.c. with water. Place solution 1 (carbonate-citrate solution) in a large beaker and add the copper solution slowly with constant stirring. The mixed solution is stored in a rubber-stoppered bottle. It does not deteriorate on long standing.

*Procedure.*—To 500 of Benedict's solution, in a test-tube, add eight drops of the urine to be tested and boil for three minutes. Allow to cool *spontaneously* (do not cool under tap). If glucose is present a large precipitate will form, otherwise the solution will remain perfectly clear or only slightly turbid.

**Quantitative Estimation of Dextrose.**—**Benedict's Method.**—Benedict's quantitative solution is made up as follows: Dissolve in 750 c.c. hot distilled water the following: crystals of sodium carbonate, 200 grams (or anhydrous 100); sodium or potassium citrate 200 grams, and 125 grams potassium sulphocyanate. When all is dissolved filter if the solution is not clear.

Dissolve exactly 18 grams pulverized, pure crystals of copper sulphate in 100 c.c. water and pour slowly into solution No. 1, with constant stirring. Allow to cool, add 5 c.c. of a 5 per cent. aqueous solution of potassium ferrocyanide solution and make up to 1000 c.c. with distilled water. This solution keeps indefinitely in a rubber-stoppered bottle.

**Procedure.**—If the urine<sup>1</sup> to be titrated shows very heavy reduction (*i. e.*, contains probably a large amount

<sup>1</sup> In case chloroform was used for preservation, a portion of the urine should be brought to boiling-point and quickly cooled. This rids it of the chloroform.

of sugar) dilute 10 c.c. of it to 100 c.c. with distilled water. Fill a clean and dry 50 c.c. burette with the solution.

Exactly 25 c.c. of Benedict's solution is measured with a pipette into a porcelain evaporating dish, 25 cm. in diameter, and to it about 15 grams of crystallized sodium carbonate (or 8 grams dry  $\text{Na}_2\text{CO}_3$ ) are added. Half a teaspoonful of powdered pumice or talc is added and the dish heated to boiling over a free flame until the carbonate has entirely dissolved.

The diluted urine is now run from the burette rather rapidly until a chalk-white precipitate forms, and the blue color of the mixture begins to lessen perceptibly. Now the diluted urine must be dropped from the burette very slowly—a few drops at a time—until the disappearance of the last trace of blue color which marks the end-point. The solution must be kept boiling vigorously throughout the entire titration. If the boiling mixture begins to bump or spatter add a little distilled water to make up for the loss by evaporation.

**Calculation.**—The amount of diluted urine which was necessary to reach the end-point is found by reading the burette. This figure divided by 10 gives the equivalent amount of urine (because the burette contained urine diluted ten times). This amount of urine must contain 50 mg. (0.05 gram) of dextrose, because it requires 50 mg. dextrose to reduce 25 c.c. Benedict's solution.

Let  $x$  = amount of urine containing 50 mg. dextrose, then  $\frac{.05}{x} \times 100$  = per cent. sugar in original urine.

**Urea.—Hypobromite Method.**<sup>1</sup>—*Solution A.*—Dissolve 62.5 grams sodium bromide in 400 c.c. water. Pour into a 500 c.c. volumetric flask. Add 22 c.c. pure bromine (under hood—bromine fumes are very irritating), and make up to 500 c.c. with distilled water. Preserve in rubber-stoppered bottles.

*Solution B.*—Dissolve 125 grams sodium hydroxide in 400 c.c. distilled water. Allow to cool. Pour into a volumetric flask and make up to 500 c.c. Preserve in a rubber-stoppered bottle.

Into the open arm of the clean Doremus-Hinds ureometer add a drop or two of urine; turn the stop-cock so that the urine just fills the lumen and turn to off position. Fill the closed arm with a mixture of equal parts of solutions *A* and *B*, and tilt to let out any air. Set upright and fill open arm with urine, avoiding bubbles of air. Turn stop-cock and allow 1 c.c. urine to flow in. After twenty minutes read the amount of gas in the closed arm. Each small subdivision of gas represents 0.001 gram of urea per cubic centimeter of urine. If percentage is desired move the decimal point two places to the right. Every small division means 0.1 per cent. urea.

**Urease Method.**—*Van Slyke's Modification of Marshall's Method.*—This method is accurate and so simple that the nurse with little laboratory experience can be

<sup>1</sup> The following method for the preparation of the hypobromite solution has also been recommended. Keep on hand in a rubber-stoppered bottle a 20 per cent. solution of NaOH. When ready for a test add to 40 c.c. of this solution 1 c.c. of pure bromine and shake. This solution may be used in place of the mixture of *A* and *B*, as indicated in the text.



taught to use it. The determinations, especially the first, should be done under the direction of a physiological chemist. The details of the method are given for reference of the nurse after she has learned to apply it. The method depends upon the fact that urease, a ferment found in the soy bean, is capable of converting urea into ammonia without any loss of nitrogen. The ammonia can then be blown out of the urine into a known amount of standard acid. The acid is titrated at the end of the experiment to determine how much has been neutralized by the ammonia.

The method is rapid and of advantage on account of the fact that several analyses may be run at the same time.

**PROCEDURE.**—One-half c.c. of urine<sup>1</sup> is measured into the bottom of tube *A*. Exactly 5 c.c. of a solution containing 6 grams of  $\text{KH}_2\text{PO}_4$  per liter are then run in from a burette, and 1 c.c., accurately measured, of a 10 per cent. solution of urease<sup>2</sup> is added.

<sup>1</sup> An Ostwald pipette is used, the stem of which is a heavy walled capillary tube of only 1 mm. bore. The pipette, which should deliver in about twenty seconds, is calibrated by weight for blow-out delivery, and permits measurement with an accuracy of 0.001 c.c. The pipette is allowed to deliver with its tip against the lower part of the test-tube wall until the bulb is emptied; the remainder of the contents is then blown out.

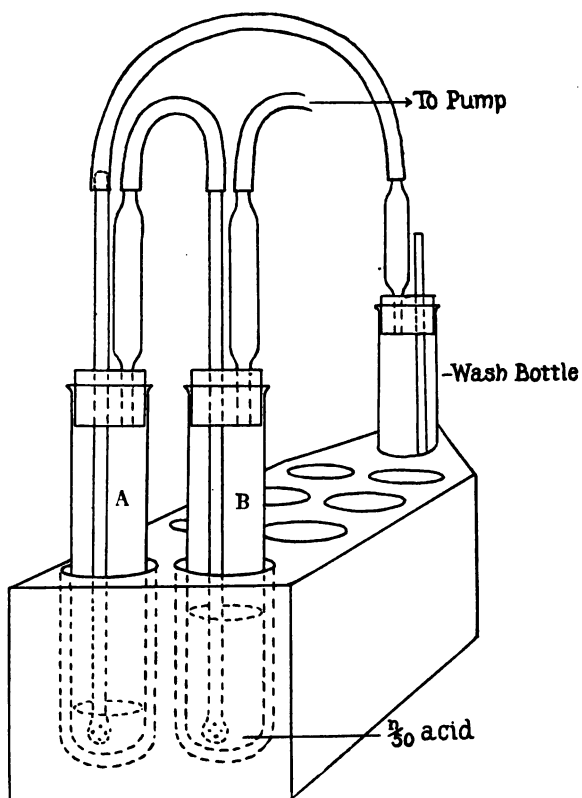
These pipettes, as well as the 100 c.c. test-tubes of special heavy glass, provided with inlet and outlet for aëration, the block holder shown in the figure, and a brass aspirator pump suitable for the method can be obtained from Emil Greiner, 45 Cliff Street, New York.

<sup>2</sup> The enzyme preparation used should be standardized as follows: A solution is made containing 3 grams, accurately weighed, of pure urea per 100 c.c. Using the special pipette described in the urine analysis, one measures into tube *A*, 0.5 c.c. of the above urea solution, 5 c.c. of 0.6 per cent.  $\text{KH}_2\text{PO}_4$ , and the amount of enzyme solution intended to be used in analysis (usually 1 c.c. of 10 per cent. enzyme). The reaction is allowed to run at room temperature (or 50° if desired)

The solutions in the tube are well mixed, 2 drops of caprylic alcohol to prevent subsequent foaming are added, and the stopper bearing the aërating tubes shown in the figure is put into place. Twenty minutes at a room temperature of 15°, or fifteen minutes at 20° or above, are allowed for complete decomposition of urea. No harm is done if the solutions are allowed to stand longer, but *the time must not be cut shorter unless more enzyme is used*. While the enzyme is acting, one measures 25 c.c. of  $\frac{n}{50}$  hydrochloric or sulphuric acid into tube *B* and connects the tubes as shown in the figure. After the time for complete decomposition of urea has elapsed the air current is passed for a half minute in order to sweep over into *B* a small amount of ammonia which has escaped into the air space of *A* during the decomposition. *A* is now opened and 4 to 5 grams of potassium carbonate measured roughly from a spoon are poured in (in order to assure most rapid removal of ammonia by air current it is necessary to have the solution at least half saturated with carbonate). The air current is now passed rapidly through the tubes until all the ammonia has been driven over into the acid in *B*. The time required for this depends on the speed of

for the length of time allowed in analysis, and the ammonia is determined as described for urine analyses. It should neutralize 25 c.c. of  $\frac{n}{50}$  acid. If it falls slightly short, it is well to repeat the test, doubling the time interval, as some samples of urea are not 100 per cent. pure, and the short figure may be the fault of the urea, not of the enzyme. If in the longer interval no more ammonia is formed than in the shorter, the urea decomposition was complete in the shorter time, and the enzyme is sufficiently active. This method is described in detail in the article by Van Slyke and Cullen, *Journal Biological Chemistry*, vol. xix, No. 2, October, 1914.

the air current. With a rapid pump or house vacuum it is possible to aërate completely in five minutes;



Apparatus for determining urea content by means of urease.  
(Journal of Biological Chemistry.)

while a slow pump may require a half hour. The time required for complete aëration is determined for the

particular vacuum used by trial, and a safe margin allowed in the determinations. When the aëration is finished the excess acid in *B* is titrated with  $\frac{n}{50}$  NaOH.

The operations can be concisely summarized in the following diagrammatic form:

1. Measure into *A* . . .  $\left\{ \begin{array}{l} 0.5 \text{ c.c. urine.} \\ 5.0 \text{ c.c. 0.6 per cent. KH}_2\text{PO}_4. \\ 1.0 \text{ c.c. 10 per cent. urease.} \\ 2 \text{ drops caprylic alcohol.} \end{array} \right.$   
 Place stopper as shown in Fig. 1 and let stand fifteen minutes.
2. Meanwhile measure into *B* . . .  $\left\{ \begin{array}{l} 25 \text{ c.c. } \frac{n}{50} \text{ acid.} \\ 1 \text{ drop 1 per cent. sodium alizarin sulphonate indicator.} \\ 1 \text{ drop caprylic alcohol.} \end{array} \right.$
3. (After 15 minutes standing) aërate one-half minute. Then open *A* and add 4 to 5 grams  $\text{K}_2\text{CO}_3$ .
4. Aërate all  $\text{NH}_3$  from *A* over into *B*.
5. Titrate excess acid in *B* with  $\frac{n}{50}$  NaOH.
6. Calculate:  $0.056 \times \text{c.c. } \frac{n}{50} \text{ acid} = \text{grams urea} + \text{ammonia nitrogen per 100 c.c. urine.}$

In order to determine the *ammonia nitrogen* alone one measures 5 c.c. of urine into *A*, adds the carbonate at once, and aërates as described above. The acid neutralized is multiplied in this case by the factor 0.0056, to give the per cent. of ammonia nitrogen. No extra time is required for the ammonia determination, as one merely aërates the extra pair of tubes in series with the same air current used for the ammonia + urea determination. As a matter of fact, one can conveniently run as many as eight pairs of tubes on the same air current, taking the precaution at the end of the aëration to disconnect the series in the middle first in order to prevent back suction.

**Bile.—Gmelin's Test.**—The urine is superimposed over nitric acid in exactly the same way as in the

test for protein. In the bile test, however, the nitric acid should be slightly yellow instead of clear. Nitric acid may be turned yellow by adding a few pine shavings (Emerson). The yellow color indicates the presence of nitrous acid,  $\text{HNO}_2$ . In the presence of bile the line of contact of the two liquids will present strata of colors varying from green, blue, violet, red, and yellow.

In urine containing bile, the foam is always yellow—otherwise it is white.

**Indican.—Jaffe's Test.**—To 5 c.c. urine in a test-tube add 5 c.c. concentrated  $\text{HCl}$ . To the mixture add 2 c.c. chloroform and 5 drops of a filtered saturated solution of bleaching powder. Shake thoroughly. A blue or red coloration of the chloroform indicates that indican was present in the urine.

**Diazo Reaction.**—*Solution A.*—5 grams of sodium nitrite in 1 liter water (this solution should not be used after standing two weeks). Preserve in glass-stoppered bottle.

*Solution B.*—5 grams of sulphanilic acid and 50 c.c.  $\text{HCl}$  in 1 liter distilled water. Preserve in glass-stoppered bottle. Mix solutions *A* 1 part to 50 parts solution *B*.

To 5 c.c. urine add 5 c.c. mixture of reagents as above indicated. Mix thoroughly by shaking, add quickly 1 c.c. ammonium hydroxide. If the fluid and the foam turn red the test is positive. On standing a precipitate is formed leaving a supernatant fluid, which is green, blue, or violet. In normal urine the reagent produces a brownish-yellow color.

# INDEX.

## A

**ACETALDEHYDE**, 157  
**Acetanilid**, 210  
**Acetic acid**, 158  
**Acetylene**, 121. *See* Calcium carbide.  
**Acids**, 106  
     acetic, 158, 183  
     amino-, 204  
     carbolic, 195  
     fatty, 185  
     formic, 146, 183  
     hydrochloric, 83  
     lactic, 169  
     organic, 146  
     propionic, 183  
     unsaturated, 185  
     uric, 245  
**Albumin in urine**, 253  
**Albumins**, 219  
**Alcohol**, butyl, 161  
     di-atomic, 162  
     ethyl, 154  
     grain, 154  
     higher, 161  
     methyl, 144  
     mon-atomic, 162  
     primary, 161  
     propyl, 161  
     secondary, 162  
     tri-atomic, 162  
     wood, 154  
**Aldose**, 170  
**Alkaloids**, 212  
**Alum**, 129  
**Aluminum**, 129  
**Amalgam**, 125

**Amino-acid**, 204  
**Ammonia**, 202  
**Ammoniated mercury**, 127  
**Ammonium acetate**, 203  
     hydroxide, 203  
**Analysis**, volumetric, 109  
**Anilin**, 209  
**Animal sugar**, 175  
**Anode**, 43, 48  
**Antimony**, 115  
**Antitoxin**, 228  
**Argyrol**, 134  
**Arsenic**, 114  
     antidote, 130  
**Arsine**, 114  
**Asymmetric carbon atom**, 168  
**Atomic weights**, 39  
**Atoms**, 30  
**Avogadro's hypothesis**, 40

## B

**BACTERIA**, nitrogen-fixing, 202  
**Baking powder**, 103  
**Bases**, 106  
**Benedict's method**, 254, 255  
**Benzaldehyde**, 198  
**Benzene**, 191, 192  
     ring, 193  
     series, 191  
**Benzol**. *See* Benzene.  
**Bichloride of mercury**, 126  
**Bile**, 261  
     test for, 262  
**Bismuth**, 116  
**Bitter almonds**, oil of, 199  
**Bleaching powder**, 86, 120

Blood, 226  
 alkalinity of, 227  
 amount of, 230  
 cells, 228  
 clot, 226  
 functions of, 231  
 salts of, 227  
 serum, 227  
 specific gravity of, 230  
 Boiling-point, 57  
 Borax, 69  
 Bromine, 89  
 Bromoform, 143  
 Bunsen burner, 121  
 Butane, 160

**C**

CALCIUM, 119  
 carbide, 139  
 Cane sugar, 173  
 Carbohydrates, 180  
 Carbon, 138  
 monoxide, 138  
 Carbonates, 138  
 Carboxyl group, 146, 168, 169  
 Casein, 235  
 Catalyzers, 78, 135  
 Cathode, 43, 48  
 Cellulose, 177  
 Centigrade, 56  
 Charcoal, 138  
 Chemical change, 23  
 Chloral, 157  
 Chlorine, 81  
 preparation of, 81  
 properties of, 82  
 uses of, 82  
 Chlormethane, 143  
 Chloroform, 143  
 Coal tar, 191  
 dyes, 210  
 Colloids, 135, 227  
 Condensation, 60  
 Conservation of mass, 32  
 Corrosive sublimate, 126  
 Creatinin, 246  
 Crenation, 230  
 Cresols, 198  
 Crystallization, 66  
 Cyanogen, 208

**D**

DEAD Sea, 81  
 Destruction of matter, 31  
 Developer, 134  
 Dextrose, 168, 173  
 in blood, 231  
 Dialysis, 135, 219, 227  
 Diamond, 138  
 Diazo reaction, 262  
 Diazonium, 211  
 Dietary, carbohydrates in, 182  
 proteins in, 222  
 Digestion, carbohydrate, 180  
 intestinal, 180  
 mouth, 180  
 pancreatic, 180  
 stomach, 180  
 Disaccharid, 175  
 Distillation, 66  
 fractional, 67

**E**

ELEMENTS, 27  
 Empirical formula, 142  
 Energy, 32, 181  
 kinetic, 33  
 latent, 33  
 Epsom salt, 123, 124  
 Ethers, 150  
 Evaporation, 60

**F**

FAHRENHEIT, 56  
 Fats, 183  
 body, 185  
 digestion of, 187  
 milk, 236  
 vegetable, 186  
 Fehling's solution, 170, 173, 253  
 Fermentation, 156  
 Ferments, 156, 159  
 in blood, 231  
 oxydizing, 231  
 Fibrin, 226, 232  
 Filter, porcelain, 70  
 sand, 70

Flame test, 121, 123  
 Fluorine, 93  
 Fluorite, 119  
 Food value of carbohydrate, 182  
   of fats, 187  
   of milk, 237  
   of proteins, 221  
 Fool's gold, 95, 114  
 Formaldehyde, 144  
 Formalin, 144  
 Formic acid, 146  
 Formula, graphic, 169  
 Fowler's solution, 115  
 Fractional distillation, 67  
 Freezing, 61

**G**

GASOLINE, 161  
 Globulins, 218  
 Glucose, 167, 173  
 Glycerine, 162, 184  
 Glycogen, 181  
 Glycoproteids, 220  
 Gmelin's test, 261  
 Gram, 24  
 Grape sugar, 166  
 Graphite, 138

**H**

HARDNESS of water, permanent,  
   69  
   temporary, 69  
 Heat, 60  
   of condensation, 60  
   of evaporation, 60  
   of freezing, 62  
   of melting, 62  
 Heller's test, 253  
 Hemolysis, 230  
 Hexane, 171  
 Hexose, 170  
 Homogenized milk, 236  
 Honey, 170  
 Hydrated lime, 21  
 Hydrocarbons, saturated, 160,  
   163  
 Hydrogen, 48

Hydrogen, nascent, 49  
   peroxide, 77  
   properties of, 49  
   uses of, 49  
 Hydrolysis, 174  
 Hydrometer, 55  
 Hydroxyl, 73  
 Hygroscopic, 121, 162  
 "Hypo," 134  
 Hypobromite solution, 257

**I**

INCOMPATIBLES, 75  
 Indican, 245  
   test for, 262  
 Indicators, 108  
 Indigo, 209, 210  
 Inversion, 174  
 Invert sugar, 174  
 Invertase, 174  
 Iodine, 90  
   preparation of, 91  
   properties of, 92  
 Iodoform, 143  
 Ionization, 73  
 Iron, 129  
 Isotonic, 229

**J**

JAFFE'S test, 262

**K**

KETOSE, 170  
 Kjeldahl, 225

**L**

LACTASE, 175  
 Lactose, 175, 236  
 Laking of blood, 230  
 Law of combining weights, 36  
   of conservation of mass, 32  
   of constant proportions, 35  
   of multiple proportions, 35  
 Lead, 133



Levulose, 170, 173  
 Light, polarized, 167  
 Lime, 19  
 Lipase, 187  
 Liquor chlori compositus, 82  
 Lithium, 110  
 Litmus, 85, 108  
 Lunar caustic, 133  
 Lysis, 174

**M**

MAGNESIA, 124  
 Magnesium, 123  
 Malt sugar, 156, 175  
 Manganese, 131  
 Mannite, 166  
 Marsh gas, 141  
   series, 153  
   test, 115  
 Marshall's method, 257  
 Melting, 62  
 Mercury, 125  
 Meta compounds, 195  
 Metals, 25  
 Methane, 142  
 Methyl alcohol, 144  
 Milk, 235  
   clot, 235  
   cow's, 237  
   human, 237  
   salts, 236  
   sugar, 175, 236  
 Mine, damp, 141  
 Moisson, 138  
 Molecular solutions, 106  
 Molecules, 30  
 Monosaccharid, 170

**N**

NITRIC acid test for albumin,  
 253  
 Nitro benzene, 193  
 Nitrogen, 202

**O**

OIL of bitter almonds, 199  
 corn, 186

Oil, cottonseed, 186  
   olive, 186  
   peanut, 186  
 Olein, 185  
 Ortho compounds, 195  
 Osmosis, 228  
 Oxidation, 27  
 Oxides, 45  
 Oxygen, 26, 43  
   preparation of, 43  
   properties of, 44  
   uses of, 44  
 Ozone, 46

**P**

PALMITIN, 185  
 Para compounds, 195  
 Paraffins, 160  
 Paraldehyde, 157  
 Pentane, 160  
 Pentoses, 172  
 Petroleum, 160  
 Phenol, 195  
 Phenylhydrazin, 210  
 Phosphorus, 112  
   effects of, 113  
   forms of, 113  
   uses of, 113  
 Photography, 134  
 Physical change, 22  
 Plasma, 226  
 Plaster of Paris, 120  
 Platinum, 134  
 Polariscope, 167  
 Polysaccharid, 173  
 Potassium, 109  
   permanganate, 131  
   sulphocyanate, 209  
 Pressure, osmotic, 229  
 Pyridin, 213  
 Priestly, 27  
 Propane, 158, 160  
 Proportions, constant, 35  
   multiple, 35  
 Proteids, 218  
   compound, 219  
   derived, 220  
   glycoproteids, 220  
   simple, 218  
 Proteins, 216

Proteins, classification of, 218  
 digestion of, 220  
 occurrence of, 217  
 Pro-thrombin, 232  
 Purification of substances, 66  
 Putrefaction, 156

**Q**

QUICK-LIME, 19

**R**

RADICAL, organic, 144  
 Reducing power, 169  
 Reversible reaction, 51, 119  
 Rusting process, 25

**S**

SACCHAROSE, 173  
 Saline, 101, 230  
 Salt, 102  
 Saltpetre, 110  
 Salvarsan, 115  
 Silver, 133  
 Soda, 103  
   lye, 104  
 Sodium, 100  
   chloride, 102  
   hydroxide, 104  
 Solubility, effect of temperature on, 65  
 Solutions, 64  
   Fehling's, 170, 173  
   hypertonic, 229  
   hypotonic, 229  
 Specific gravity, 55  
   of urine, 240, 250  
 Spectroscope, 124  
 Spectrum, 124  
 Springs, 68  
 Starch, 176  
 Stearin, 185  
 Strontium, 121  
 Structural formula, 142  
 Sugars, 165  
   of lead, 133

Sugars, reducing of, in urine,  
   253  
     Benedict's method, 254  
     Fehling's method, 253  
 Sulphocyanates, 209  
 Sulphur, 95  
   dioxide, 97  
 Symbols, 28

**T**

TARTAR emetic, 116  
 Thermometer, 56  
 Thrombin, 232  
 Toluene, 197  
 Toluol. *See* Toluene.  
 Tonicity, 229

**U**

URANALYSIS, 249  
 Urea, 243  
   constitution of, 244  
   estimation of, 257  
     hypobromite method, 257  
     urease method, 257  
     Van Slyke's method, 257  
 Urease method, 257  
 Ureometer, 257  
 Uric acid, 245  
 Urine, 240  
   acidity of, 252  
   albumin in, 253  
   amount of, 240, 250  
   collection of, 249  
   color of, 241, 251  
   odor of, 241, 252  
   organic bodies in, 243  
   preservation of, 249  
   reaction of, 242, 252  
   salts in, 242  
   specific gravity of, 240, 250  
   total solids in, 251  
 Urochrome, 241  
 Uronoid, 241

**V**

VALENCE, 97  
 Van Slyke method, 257

Vinegar, 158  
  method of, 158  
Vitamines, 213  
Volumetric analysis, 109

**W**

WATER, 53  
  composition of, 72  
  hard, 69  
  potable, 70  
  properties of, 54  
  rain, 70  
  river, 68  
  sea, 68  
  soft, 69

Water, as standard of comparison, 55  
  synthesis of, 72  
  uses of, 54  
Weights, 24  
Wood, 177

**X**

XANTHOPROTEIC reaction, 218  
X-rays, 117  
Xylene, 197  
Xylol. *See* Xylene.

**Y**

YEAST, 155, 174

OCT 12 1915

